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# LABORATORY EXERCISES IN CHEMISTRY.

BY

EDWIN J. BARTLETT.



- CHAPTER I. Object Lessons for Beginners.  
CHAPTER II. The Ordinary Calculations.  
CHAPTER III. Experiments in Weight and Volume Relations.  
CHAPTER IV. Suggestions How to Make Chemicals.  
CHAPTER V. Words and their Uses.



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BOSTON.

NEW YORK.

CHICAGO.

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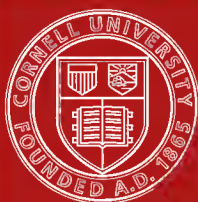
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# LABORATORY EXERCISES

IN

# CHEMISTRY

BY

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LEACH, SHEWELL, & SANBORN,

BOSTON, NEW YORK, CHICAGO.

A, 146797

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## PREFACE.

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THIS book is a Laboratory Guide for beginners in chemistry, and something more.

Part I is the common ground over which beginners must travel. It has been many times worked over with reference to clearness of instructions; concentration of attention upon the essential thing by avoiding complicated and distracting apparatus; the elimination of dangerous or seriously irritating experiments; economy in chemicals; relation to subsequent work.

It is thought especially adapted to conditions which are quite prevalent—the necessity of economy in supplies and teaching force. The numerous references are essential and call the pupil's attention to intersection of properties.

Part II contains the arithmetical guide, absolutely necessary in some form for everything beyond the most elementary stage. It tries to draw a careful distinction between those computations which are original in their nature and those which merely utilize accumulated facts.

If the pupil can solve the limited number of diverse illustrative problems given, he will probably have no difficulty in the practical cases which will arise. If the teacher prefers, he can easily multiply the simpler problems. It should be remembered that here arithmetic is a means, not an end.

Part III contains a series of quantitative experiments, which illustrate fundamental principles. They are fairly accurate and do not require complicated or costly apparatus, nor anything outside of the rather familiar forms. It has been thought best to give full working directions, but to leave explanations for the pupil's thought with the guidance of the teacher.

Part IV is simple enough to those who have learned what it

contains; the instructions are merely those which every novice in chemistry must have before he can pass to the slightest degree of independence. Perhaps this Part will not be used as a distinct course of study in many schools, but rather as containing convenient directions to be utilized from time to time.

Part V contains a list, not exactly of definitions, but of terms and their uses. They are not to be memorized, but learned in the proper applications.

A large portion of Part II is essential to III; but after the arithmetic is applied Part III may be combined with I; though it seems to the writer preferable to have phenomena and manipulation somewhat familiar first.

The time in the Laboratory (of course approximate) is Part I, without bracketed experiments, 30 hours; bracketed experiments, 10 hours; Part III, 30 hours; Part IV, 10 or more hours.

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It will be seen that this book contemplates giving the beginner the absolute essentials, — language, notation, substances, processes, qualities, quantities, — all from the experimental standpoint. The desirable but not immediately essential matters, — history, sources, manufactures, applications, etc., — can easily be given by oral, text, or reference instruction according to time available and according to the pupil's interest.

The Atomic Theory should be treated when it can be understood as an explanation of facts and laws, not as a dogma to which all facts and laws must correspond. It is a well-fitting glove on a substantial hand, but is not the hand.

A list of the chemicals and apparatus needed by the pupil for the work laid out in the book is given; but all the things desirable and even necessary in a laboratory are not given.

The full list of materials for each day's work the teacher will find in the paragraphs for the day, but for his convenience in ordering goods and preparing reagents a list is given in order on p. 10. This list does not repeat articles once provided for,







## INTRODUCTION.

---

THE laboratory study of chemistry involves the study of substances, of processes, of weight and energy relations. The first two are for the last, and are in their nature precedent.

The absolute essentials for the pupil's success are neatness, orderliness, common sense, the ability to form a plan from printed instructions and to adapt means to ends in carrying it out.

The pupil should always call things by their right names; and think of them, at any rate, by their distinguishing names. There is a whole world of difference between saying and thinking "that thing" and saying and thinking "a 60cm. royal Berlin porcelain evaporating dish, round bottomed, lipped."

The duplication and repetition of experiments have great value. By duplication the student tests his own work and learns, too, of its regular, and not accidental, character; by repetition he corrects his mistakes of unfamiliarity, and uses his own mind for improvement.

It is indispensable that scientific notes should be taken *at the time of the observations*; it is desirable that they should be put into good form later.

All chemical properties are correlative.  $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$ . It is equally a property of zinc to decompose

sulfuric acid, and of sulfuric acid to give up its hydrogen for zinc.

Every chemical action is dependent upon a certain set of conditions. Success is a mere matter of acquired skill within narrow lines. The interpretation and correction of failure is far broader.

Processes in chemistry are direct inferences from the properties, chemical or physical, of the substances involved.

The older chemist enlarges his knowledge and increases his skill

1. By his own investigations, often — perhaps most often — upon seeming trifles.

2. By reference to the great stock of accumulated knowledge — text-books and monographs.

3. By noting the fresh results of others' work — in the journals.

4. By taking numerical constants, except a few familiar ones, from their special tables when he wants them.

The beginner should use the same methods.

### *Supplies.*

In purchasing supplies for the laboratory it is to be taken for granted that the work will continue, and material should be obtained with reference to permanency of stock. It is usually very poor economy to purchase little scraps in ounce or half-ounce bottles.

Most of the chemicals, with reasonable care, last until they are used. A few need special care. Strong nitric acid should be kept from the light; lime, lime water, the strong alkalis, all deliquescent and efflorescent salts, bisulfid of





carbon, iodine, bromine, and all volatile substances must be kept in tightly closed bottles. The concentrated acids are purchased and diluted for use. Salts are purchased and dissolved in water as needed; ordinary phosphorus is always kept in water. Sodium and potassium are kept in naphtha, or even kerosene.

Many of the articles in the following list are usually obtainable near by, as the teacher will see upon looking it over. Sugar, starch, lime, threepenny nails, No. 28 iron wire, salt, bleaching powder, washing soda, etc., are usually not far away.

As a source of heat, alcohol lamps are satisfactory except in a few cases, as bending glass. When Bunsen burners are used for heat, they must be watched very closely, as there is a great tendency in the novice to push a burner under something and let it go, while he proceeds to consider something in another part of the room.

It is both impracticable and undesirable to furnish each pupil with any considerable individual stock of chemicals for this particular work. The chemicals are best arranged for the daily work in definite order upon a common table. Dilute acids and ammonia may be on the pupil's table.

### *The Apparatus.*

The plan of these exercises contemplates the following outfit for each pupil, which the teacher can modify to suit himself in accordance with his experience and preference. This is not intended to be a complete stock for a laboratory.

1 test tube rack.

12 test tubes 4" or 6"x $\frac{5}{8}$ ".

2 beakers, plain, about 50 and 75 c.c.

1 evaporating dish, about 60 c.c.

- 2 side-necked test tubes, 6''x1''.
- 1 flat-bottomed Erlenmeyer flask, 200-250c.c.
- 1 thistle tube.
- 2 gas cylinders, 300c.c., ground edges.
- 2 cylinder covers, plates of glass 3'' square cut from scrap window glass.
- 1 funnel 2½''. 1 clay crucible, 2 oz.
- 1 piece of sheet iron, or iron netting 3'' square.
- 1 piece galvanized or tinned iron 6''x¾'', bent to a V for jar support.
- 1 glass rod and spatula, 5'' long, pointed at one end, flat at the other, made from ¼'' or ⅜'' tubing.
- 1 delivery tube, 15'' of ¼'' tubing.
- 1 low 2 qt. pan of earthenware or tin.
- 1 iron ring stand.
- 1 pair test tube holders.
- 1 alcohol lamp, 4 oz., globe-shaped.
- Rubber connections, corks to fit test tubes, side-necked test tubes and flask, glass tubing as required ⅜'', ¼'', ⅜'', ⅝''.

There will be needed occasionally in the pupil's work 50 or 100c.c. graduates, triangular files, Bunsen burners, cork borers, mortars and pestles, chemical thermometers, scales, blast-lamp.

The materials needed in Parts III and IV are given there.

The chemicals for Part I arranged in order of use by paragraphs. Those once given are not mentioned a second time. The proportions are satisfactory for the purposes of this book; not for all other purposes.

- |                                    |                                     |
|------------------------------------|-------------------------------------|
| 2. scrap glass tubing.             | 6. mercuric oxid.                   |
| 3. crystallized copper sulfate.    | splinters of straight-grained back- |
| ferric oxid.                       | ing.                                |
| 4. dil. hydrochloric acid (1:3).   | zinc dust.                          |
| dil. ammonia (1:2).                | flowers of sulfur.                  |
| litmus paper, red and blue, cut in | crystallized barium chlorid.        |
| narrow strips 1½'' long.           | 7. granulated zinc, made by melting |
| 3 or 4 other convenient reagents.  | common scrap in an iron ladle       |
| 5. broken roll sulfur.             | and pouring slowly into water.      |
| strips of copper foil.             | conc. commercial sulphuric acid.    |
| dil. nitric acid (1:3).            | conc. commercial hydrochloric       |
| 3-penny nails.                     | acid.                               |







8. potassium chlorate, fine, ground,  
if necessary, in a *clean* mortar.  
sol. of silver nitrate (1 : 40 distilled  
water).  
distilled water.  
manganese dioxid, powdered.  
iron wire (on spools, No. 28).  
barium peroxid.  
dil. starch paste. Shake two or  
three lumps of common starch  
with 10c.c. of cold water, boil  
till clear, add to about 200c.c.  
of water.  
sol. of iodid of potassium (1 : 40).
9. granulated sugar.  
cryst. carbonate of soda.  
quicklime in small pieces (par-  
tially air-slaked will do).  
lead oxid (litharge).  
arsenious oxid ("white arsenic").  
"fused" or anhydrous calcic  
chlorid.
10. calcium sulfate (plaster of Paris).  
hyposulfite of soda.  
common table or dairy salt.  
ordinary saltpeter.
11. hydrogen peroxid (ord. "12 to 15  
vol.") may be diluted with an  
equal part of water.  
sol. of potassium permanganate  
(1 : 50).  
fresh lead sulfid, made by  
precipitating lead nitrate or  
acetate with either hydrogen  
sulfid or ammonium sulfid and  
washing 3 or 4 times by decan-  
tation.
12. paraffin (cut in lumps about  
chestnut size).  
hydrofluoric acid (in ceresine  
bottle).
- scrap window glass (cut about  
2" x 3").  
filter paper, common, 9.c.m. to  
11c.m.
13. "chlorid of lime" or bleaching  
powder, obtainable in paste-  
board boxes.  
printed paper.  
stale urine or other decomposing  
material diluted with water, or  
H<sub>2</sub>S water.
14. rock salt, broken.  
strong ammonia (with great cau-  
tion, for this experiment only).
15. small or broken crystals of potas-  
sium bromid.  
bromin, about ½ an inch in a  
bottle, in the hood, *must* be  
taken with a dropping tube and  
not poured.  
ferrous sulfid, broken into little  
lumps.
16. small or broken crystals of potas-  
sium iodid.  
iodin crystals.  
common alcohol.  
bisulfid of carbon.  
sol. of KI (same as 8).  
baking powder.
17. 2 oz. hessian crucible.  
ferrous sulfid (powdered and sifted).  
"dry" carbonate of soda.  
selected charcoal (about egg size).
18. lead nitrate or acetate sol. (1 : 40  
made clear by a few drops of  
nitric or acetic acid).  
antimony chlorid (1 : 40 water  
acidulated with HCl).  
ferrous chlorid (1 : 40, it may be  
made by dissolving a nail in  
conc. HCl).

19. sheet copper (cut in strips about  $1\frac{1}{2}'' \times \frac{1}{2}''$ ).  
extract of logwood (enough added to water to give a good color).  
permanganate of  $\text{KI}$  diluted with 3 or 4 parts of water.  
sulfite of soda.
20. sol. of barium chlorid ( $1:20$ ).  
sol. of sodium sulfate ( $1:20$ ).  
sol. of di-sodium phosphate ( $1:20$ ).
21. ordinary ammonium chlorid.  
commercial potassium nitrite.
22. ordinary phosphorus (cut into pieces under water, and kept under water till used).  
lime water, see Part IV.  
common cotton batting.  
any of the commercial extracts of beef,  $\frac{1}{2}$  teaspoonful to 1 qt. of water.
23. sol. of ammonium chlorid or ammonium sulfate (about  $1:20$ ).  
 $\text{NaOH}$  sol. ( $1:10$ )
24. sols. of stannous and cadmium chlorids ( $1:40$ , fresh. may be made by dissolving the metals in conc.  $\text{HCl}$ ).  
ammonium sulfid made by *saturating* dil. ammonia with  $\text{H}_2\text{S}$  gas and then adding as much more ammonia.
25. ammonium nitrate cryst.
26. nitric acid, sp. gr. 1.2.
28. stick tin is melted in a Bunsen flame, the drops fall upon a smooth surface like slate, spread out thin and are cut in strips.  
[horse hair.]
29. [red or amorphous phosphorus.]
30. [calcium phosphid.]
31. sol. of  $\text{CuSO}_4$   $1:20$ .  
hypophosphorous acid (sp.gr. 1.15; dilute with an equal volume of water).
32. [glacial phosphoric acid.]
33. charcoal powdered and sifted.
34. metallic antimony broken in mortar.  
natural sulfid of antimony, powdered.
35. platinum, or iron wire.  
borax "dry."
36. silicate of soda, soluble glass (the commercial article in two or three times its volume of water).  
[fine sea sand.  
powdered fluor spar.  
quicksilver.]
37. powdered bone black.
38. formic acid (sp.gr. 1.06 will be suitable).
39. carbonate of lime (marble dust).  
turpentine.
40. caoutchouc (crude rubber gum, in lumps as large as a pea).
41. lycopodium.
42. magnesium ribbon, 3'.  
cadmium (dropped like tin - 28).  
copper carbonate.  
mercuric nitrate.  
ferrous oxalate.  
ferric chlorid sol.  $1-20$ .
46. lead carbonate.  
[ferrous chlorid.]
48. ammonium carbonate sol.  $1-10$ .
49. sodium carbonate sol.  $1-20$ .
50. sol. of mercuric chlorid  $1-40$ .  
sol. of copper nitrate  $1-20$ .  
strips cut from sheet lead.  
potassium ferricyanid sol.  $1:20$ .





***Chemical Notation.***

The chemical notation is merely a convenient system of abbreviations expressing substances of chemical purity and some essential facts about them.

The **symbol** of an element is an abbreviation of its Latin name, using the initial letter, or the initial and some characteristic subsequent letter. See Table, p. 61.

Generally the Latin initial agrees with the English. The following are the more important exceptions : —

<b>Na.</b> (Natrium) for Sodium.	<b>Pb.</b> (Plumbum) for Lead.
<b>K.</b> (Kalium) for Potassium.	<b>Fe.</b> (Ferrum) for Iron.
<b>Sb.</b> (Stibium) for Antimony.	<b>Hg.</b> (Hydrargyrum) for Mercury.
<b>Ag.</b> (Argentum) for Silver.	<b>Sn.</b> (Stannum) for Tin.
<b>Au.</b> (Aurum) for Gold.	<b>Cu.</b> (Cuprum) for Copper.

In all accurate chemical notation the symbol means a certain definite proportion of the element called its atomic weight. Thus O means 16 parts of oxygen and O<sub>2</sub> 32 parts.

The symbol is, however, often employed to denote the substance, e. g. "Put a glowing splinter into the jar of O."

A **formula** represents the composition of a chemical compound. It indicates what elements make up the compound *and the proportion of each element*. Thus NaCl means a substance consisting of 23 parts of sodium and 35.4 of chlorine. HNaCO<sub>3</sub> consists of 1 part hydrogen, 23 parts sodium, 12 parts carbon, and 48 oxygen. The whole is 84; so the elements are  $\frac{1+23+12+48}{84}$ .

In the case of gases and volatile liquids the formula represents the weight of definite standard volumes, usually called "two volumes," or "the molecular volume." Thus HCl, H<sub>2</sub>S, H<sub>3</sub>N, H<sub>4</sub>C — all gases — would represent equal volumes in the same discussion.

In writing a formula the order of the elements is usually

not important — KOH and KHO are the same ; but in compounds of two elements it is customary to write the more metallic first — NaCl, not ClNa ; and in compounds of three elements the usual order is metallic, non-metallic, oxygen —  $\text{KNO}_3$ , not  $\text{NKO}_3$ , or  $\text{O}_3\text{NK}$ .

A chemical formula implies that the composition of the substance has been carefully determined by some one. For the methods by which formulas are determined see Part II.

A **chemical equation** is a statement of a chemical change, including all the substances that enter into and come from the change *in their relative proportions by weight* ; also, when they are gases, in their relative proportions by measure.

$\text{Zn} + 2\text{HCl} = \text{ZnCl}_2 + \text{H}_2$  means that 65 parts of zinc displace the hydrogen from 72.8 parts of hydrochloric acid, giving 135.8 of zinc chlorid and 2 of hydrogen. These numbers are relative, but attaching a definite meaning to one fixes it for all ; if 65 grammes, ounces, or pounds of zinc are used, the other numbers represent the same denomination. The numbers come from the Table of Atomic Weights, p. 61, and methods of forming and using equations are given in Part II.

The pupil should remember that a chemical equation must have been worked out experimentally by some one, or else it is only a theory. Many of the equations in the books are theories.

Many well established equations are given through Part I, in such form that the pupil can complete them correctly.

The pupil should also remember that if an equation is true, it is true for all the substances in it. Thus the equation given above may teach the behavior of zinc (with HCl), or of HCl (with zinc), or it may apply to the making of H or of zinc chlorid.

Familiarity with chemical formulas and equations is much







like the spelling and meaning of words—to be acquired by deliberate study, by constant use, and by looking up special cases as they occur. While indispensable to sound knowledge of chemistry, the mere knowledge of these symbols is of no value as an end in itself.

The Chemical Notation is the means of expression of two somewhat distinct ideas.

1. Weight and volume relations.
2. The conceptions of the Atomic Theory.

1. Certain facts, the result of experiment, which may be looked upon as independent of any theory, find expression in chemical formulas and equations. Thus  $\text{NaCl}$  expresses what may be considered permanent facts, viz. that salt is  $\frac{23}{58.4}$  sodium and  $\frac{35.4}{58.4}$  chlorin.

The mode of expressing these facts might at some time change.  $\text{Na}=23$ , because that is now thought upon sound evidence to be the comparative weight of its chemical unit. If at some time it should appear that the chemical unit of sodium is 11.5, then the composition of salt would be expressed by  $\text{Na}_2\text{Cl}$ . The composition would not change, but the mode of expression would.

2. The prevailing and seemingly well-founded views of matter known as the Atomic Theory are also carried into the formulas, and we frequently say: “ $\text{NaCl}$  means a molecule of the substance called salt, consisting of 1 atom of sodium and 1 atom of chlorin.” The evidence upon which this *explanation* of chemical facts is based must come later.

### *Chemical Names.*

Chemistry employs a system of names largely its own.

Each *Element* has a name based on: the old common name, e. g. ferrum, iron; some property, e. g. iodine, violet-like;

on some place or thing related, e. g. scandium, Scandinavia; calcium, calx, lime; on some myth, e. g. Thorium, Thor; cerium, Ceres; niobium, Niobe; on some whimsical name, e. g. nickel, kupfer-nickel, Old Nick's copper, etc.

Each **Compound** has a name which is an attempt to express its composition and which, therefore, implies in some way the names of the elements making it and their proportions. Compounds of two elements have a name ending in *id* (*ide*). NaCl—sodium chlorid, the more metal-like or electro-positive element being named first and *id* being affixed to the less metal-like.

When the same pair of elements enter into combination in two different proportions (as  $\text{SO}_2$  and  $\text{SO}_3$ ) the terminations *ous* and *ic* are often employed; both are sulfur oxids; the first *sulfurous* oxid, the second *sulfuric* oxid. Numerals are also conveniently used, sulfur *dioxid*, sulfur *trioxid*.

In compounds of three elements *ite* and *ate* are commonly used;  $\text{KNO}_2$ , potassium nitrite;  $\text{KNO}_3$ , potassium nitrate. In a few compounds this system is carried farther by the use of *hypo* and *per* as prefixes; *hypo* . . . *ite*, less oxygen than *ite*, *per* . . . *ate*, more oxygen than *ate*; ( $\text{KClO}$ , potassium hypochlorite;  $\text{KClO}_2$ , potassium chlorite;  $\text{KClO}_3$ , potassium chlorate;  $\text{KClO}_4$ , potassium perchlorate.) See Part V.

**Classes of compounds** are grouped under common names. Thus oxids, chlorids, sulfids, etc., are compounds of two elements, one of which is oxygen, chlorin, sulfur, respectively. Nitrates are compounds containing the group  $\text{NO}_3$  ( $\text{KNO}_3$ ,  $\text{NaNO}_3$ ;  $\text{Cu}(\text{NO}_3)_2$ , etc.). Sulfates contain the group  $\text{SO}_4$  ( $\text{Na}_2\text{SO}_4$ ,  $\text{CaSO}_4$ ,  $\text{Al}_2(\text{SO}_4)_3$ , etc.).

**Processes** have names, e. g. Filtration, Evaporation, Distillation, etc.

**Pieces of Apparatus** have names, and in the ordinary labora-





tory supplies are very accurately designated either by maker's number, or better by dimensions or contents. Some good apparatus catalogue is very useful in the laboratory.

*To the Pupil.*

Think.

Read each paragraph through before you begin it; and form a plan.

Wear an apron or other suitable protection to the clothes.

Arrange your apparatus in a symmetrical, orderly way, and keep it so.

Have plenty of cleaning cloths.

Put no liquids on anything too hot to hold in the hand.

When mixing strange chemicals make it a habit to be cautious; add a drop first.

Use your good sense.

Keep your notes parallel with your work, and look up all the references.

Learn formulas as you would words, in spelling and meaning. Do not think there is any sound rule for "writing formulas"; you will be much helped after a time by analogy and valence; but formulas are facts, not guesses.

You cannot *make* a chemical equation till you are well advanced in knowledge and experimental skill; but you must begin at once to read. The equations given are meant for reading lessons, with a little call upon your attention to supply missing parts.

Learn the language.

Find out what is in this book and then go beyond it for wider information.

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# LABORATORY EXERCISES IN CHEMISTRY.

## PART I.

### LABORATORY EXERCISES.

NOTE. — Experiments in brackets should always be omitted until the instructor has considered their suitability to his classes, laboratory, and other conditions.

#### *Preliminary Manipulations.*

1. Measure in cubic centimeters, and record the capacity of beakers, test tubes, evaporating dish, etc. Consider whether it is best to fill the beaker and pour into the graduated cylinder, or the reverse.

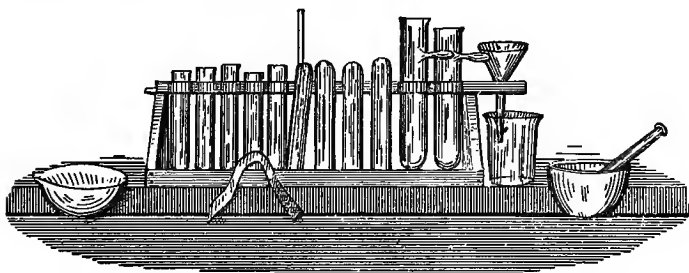


Fig. 1.

2. Cut, bend to a right angle, draw out, and close pieces of glass tubing. This should be demonstrated by the teacher, and practised upon scrap glass.

3. *Solution, Suspension, Filtration, Evaporation.* Half-fill two test tubes with water; into one put a crystal of sulfate

of copper; into the other about the same bulk of ferric oxid. Shake and warm each tube. One substance finally dissolves, giving a blue liquid; the other gives a turbid mixture. Now pour the contents of both tubes into a beaker together. Prepare a filter (demonstrate), and pour all upon it, catching the liquid which runs through — called filtrate — in evaporating dish. Evaporate the contents of the dish nearly, not quite, to dryness, by placing it over the lamp upon the ring stand. How have you separated water, oxid of iron, and sulfate of copper?

**4. Testing reaction.** 1. Half-fill two test tubes with water; into one put 2 or 3 drops of an acid (e. g. vinegar, acetic acid), into the other put 2 or 3 drops of an alkali (e. g. common ammonia, ammoniac hydrate); into each tube dip a strip of red litmus paper and a strip of blue litmus paper; note results; now exchange the papers; note results. An acid and an alkali have what effect? 2. Now take any four liquids in your reagent bottles and classify them into *acid*, *alkaline*, and *neutral* by means of litmus paper.

#### ***Physical and Chemical Properties and Action.***

**5.** 1. Examine a piece of sulfur and a strip of copper foil; write all the difference in properties you can discover by examination alone. You will note in this way only physical properties. 2. Now cautiously heat a little of the sulfur in a test tube until it boils, and then, while it is boiling, drop half the strip of copper to the bottom of the tube. Keep it boiling a minute longer. Now examine the strip; how has it changed? What is it? 3. Put the rest of the copper into dilute nitric acid in a test tube and warm till it disappears. Where is the copper? 4. Slide an iron nail into





the tube, pour off the liquid at once, and rinse the nail ; what is on it? Which has the stronger attraction for nitric acid, copper or iron? What are some of the chemical properties of copper?

***Further Illustrations of Chemical Action.***

**6.** Weigh the materials for 1 and 2. 1. Heat  $\frac{1}{2}$  gram of mercuric oxid in a clean, dry test tube. Make it very hot, but move the tube about in the flame to avoid melting the glass. Pass nearly to the bottom of the tube a splinter with a spark upon the end. It will burst into flame ; if not at the first trial, after removal and more heat. Now examine carefully the inside of the tube. Was mercuric oxid a simple or compound substance? What is the nature of the two products? This change illustrates *decomposition*. 2. Mix thoroughly 1 gr. of zinc dust and  $\frac{1}{2}$  gr. of flowers of sulfur ; (how do those weights stand related to the relative "atomic weights" of zinc and sulfur?) place the mixture upon a piece of iron (the ring stand) and apply a lighted match. Is the result simple or compound? This change illustrates *chemical union* or *combination*. 3. Weigh 1 gm. each of crystals of barium chlorid and copper sulfate ; dissolve separately in 4c.c. of water, heating the water to boiling as the substances dissolve. Now pour either into the other, drop by drop ; finally shake thoroughly ; place the test tube in the rack. After a few minutes notice that you have a clear green liquid above and a heavy white powder below. The green liquid contains copper chlorid and the white powder is barium sulfate. The metals have changed places ; this kind of change is called *double decomposition*. If you have at hand a very good filter paper, you can separate these substances, applying the knowledge of the process obtained in **3**.

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3. See whether you can get H from HCl and an iron nail. (Warm.) 4. Arrange a side-necked test tube with cork and delivery tube (Fig. 2), prepare H, using zinc, water, and HCl; let the H at first escape; then collect over water

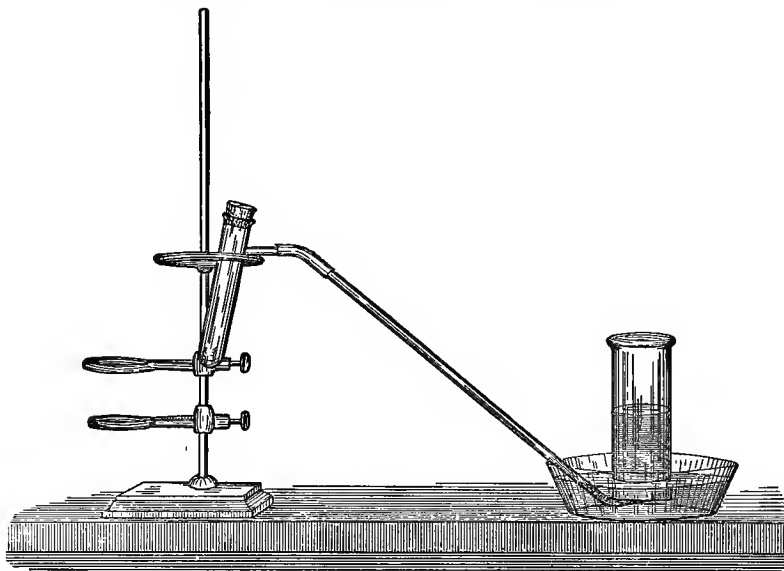


Fig. 2.

(teacher demonstrate) as follows: 5. Fill your larger beaker and, lifting it mouth downward, pour up into the other beaker, held mouth downward. Apply at once a light to each beaker; where do you find the H? Why? 6. Fill a test tube with H and light it. Fill a test tube two-thirds with air, one-third with H, light it. Notice the difference and remember never, under any circumstances, to apply a light to a hydrogen generator or any of its connections until you have satisfied yourself by collection of the gas in test

tubes that pure hydrogen is coming away. 7. Fill your gas jar with  $H$ ; lift from the tank and, holding it mouth downward, light it with a burning stick. Will the stick burn if

carried into the jar above the  $H$  flame? Can you light the stick at the  $H$  flame? What gathers on the jar inside? What properties of  $H$  do these experiments illustrate?

[8. Put in about 4 in. of  $\frac{1}{4}$  in. glass tubing a little ferric oxid, connect the tubing with the  $H$  generator which must be working well (Fig. 3). Heat the tubing very hot near the middle (see caution in 6). As the  $H$  passes over the hot ferric oxid it takes away oxygen — *reduces*. Can you discover any evidence of

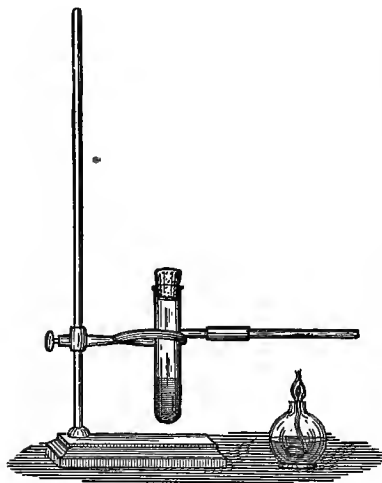
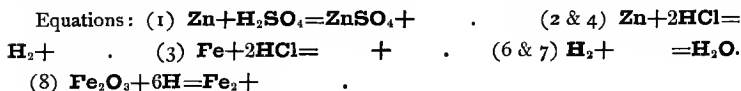


Fig. 3.

this?] 9. Evaporate some of the liquid from the side-necked test tube. You find a white solid, zinc chlorid ( $ZnCl_2$ ), which will melt if you heat too much.



### Oxygen.

8. See 6, 1. 1. Heat in a dry test tube  $\frac{1}{2}$  gm. of potassic chlorate ( $KClO_3$ ). Test the gas evolved with a glowing stick. Potassic chlorid ( $KCl$ ) is left in the tube. 2. Dissolve a portion of this  $KCl$  in distilled water; dissolve in





another portion of water a little unheated  $\text{KClO}_3$ ; now to each add a few drops of silver nitrate ( $\text{AgNO}_3$ ) solution. This is merely to show that the salt after thorough heating is different, and acts differently from the salt before heating.

3. Put in a (dry, clean) s. n. t. t. 3 or 4 gm. of  $\text{KClO}_3$  and about one-half as much manganese dioxid ( $\text{MnO}_2$ ), well mixed. Arrange completely for the collection of a gas over water, then heat the tube moderately. Remember that in all such cases the first bubbles of gas are only the air in the generator; so let a little escape and then collect two

gas jars full. Immediately take the delivery tube out of water in this and all similar cases. (Why?) Slip the cover glasses on the jars before taking them out of water, in this and all similar cases. (Why?) Take them out and let them stand covered till you are ready to use them (Fig. 4).

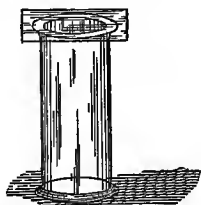


Fig. 4.

4. Into one jar (mouth upward or downward? the hydrogen density of O is 16, of air 14.44) plunge a glowing stick; when it bursts into flame remove it, blow out the flame, return it, continue this till the O is exhausted. 5. Wind a fine iron wire (No. 28) around about  $\frac{1}{5}$  of a match, using the phosphorus end (Fig. 5). Now light the match and plunge it with the wire nearly to the bottom of the other jar. A considerable portion of the wire will burn brilliantly in the jar of O. These two are typical of many cases in which combustion or burning proceeds with much more energy in pure O than it does in air, which is



Fig. 5.

only  $\frac{1}{5}$  O. 6. *Ozone* ( $\text{O}_3$ ). Put about 2c.c. of dil.  $\text{H}_2\text{SO}_4$  into a test tube; mix with it enough barium peroxid to make

a thick paste ; finally warm. This reaction produces hydrogen peroxid ( $\text{H}_2\text{O}_2$ ), oxygen, and a little ozone. Notice its sharp, irritating odor, and the effect of the gas in the tube (not the liquid) upon paper that has been dipped into a mixture of iodid of potassium (KI) solution and starch paste. (Starch and iodine form a blue compound. But starch and KI do not unless something takes possession of the potassium and releases the iodine. Ordinary oxygen cannot do this, but ozone can. Hence this reaction is a test for ozone when other substances that could take the potassium — like chlorine — are absent.)

Equations: (1)  $\text{KClO}_3 = \text{KCl} + \text{O}_2$  . (2)  $\text{KCl} + \text{AgNO}_3 = \text{AgCl} + \text{KNO}_3$ .  $\text{KClO}_3 + \text{MnO}_2 = \text{MnO}_2 + \text{O}_2$  + .

Both equations represent only the final result. (5)  $3\text{Fe} + ? \text{O} = \text{Fe}_3\text{O}_4$ .

(6)  $\text{BaO}_2 + \text{H}_2\text{SO}_4 = \text{H}_2\text{O}_2 + \text{BaSO}_4$  .  $m\text{H}_2\text{O}_2 = m\text{H}_2\text{O} + n\text{O}_2 + p\text{O}_3$  (values of m, n & p unknown).

### **Water ( $\text{H}_2\text{O}$ ).**

**9.** See **6**, 3, and **7**, 7 & 8. 1. Hold a dry cold beaker inverted around the lamp flame an instant. 2. Heat a dozen granules of sugar in a test tube ; does water appear? [3. Heat a little piece of meat in a test tube.] 4. Heat a crystal of carbonate of soda ("washing soda") in a test tube. What do 1, 2, 3, 4 show, respectively, about water? 5. Put in a side-necked tube with *clean* delivery tube 10 c.c. of water, boil with a small flame, and catch in a beaker a few c.c. of the condensed steam. It is "distilled water"; its purity depends on what was in the water at the outset. Why? You cannot get much this way. Why? See 28, 1, and Part IV, 6. Test the reaction (see **4**) of distilled water. 6. Test the reaction of water which has been shaken up with the fumes of burning sulfur (1 c.c. of water, one or two sulfur







matches with the tip cut off). 7. Test the reaction of water shaken up with a tiny lump of lime. 8. Test the reaction of water (5c.c. or so) which has been boiled with  $\frac{1}{10}$  gm. of lead oxid. 9. Test the reaction of water which has been boiled with  $\frac{1}{10}$  gm. of arsenious oxid. (Arsenious oxid is poisonous and must not be used carelessly. The reactions in 8 & 9 are slow and feeble.) 10. *Water of crystallization.* See 4. Heat in an evaporating dish a crystal of copper sulfate, holding a dry cold beaker over it. What appears on the beaker, and what do you hear and see in the dish? 11. *Efflorescence.* Take a lump of fresh washing soda, scrape off any white powder on its surface and let it stand exposed to the air. In a dry room you will see some change in ten minutes and a very great change in twenty-four hours. It parts easily with its water of crystallization. 12. *Deliquescence.* Put on a piece of glass a little "fused" (or anhydrous) calcic chlorid ( $\text{CaCl}_2$ ); see how the smallest particles soon liquefy; if, from the dryness of the air, this does not take place in a few minutes, breathe on the  $\text{CaCl}_2$ . For what can deliquescent substances be used?

Equations: (4)  $\text{Na}_2\text{CO}_3, 10\text{H}_2\text{O} = \text{Na}_2\text{CO}_3 + \quad .$  (6)  $\text{S} + \text{O}_2 = \quad .$   
 $\text{SO}_2 + \text{H}_2\text{O} = \quad .$  (7)  $\text{CaO} + \text{H}_2\text{O} = \quad .$  (8)  $\text{PbO} + \text{H}_2\text{O} = \quad .$   
 (9)  $\text{As}_2\text{O}_3 + ? \quad \text{H}_2\text{O} = ? \quad \text{H}_3\text{AsO}_3.$  (10)  $\text{CuSO}_4, 5\text{H}_2\text{O} = \text{CuSO}_4 + \quad .$

### *Solution of Salts in Water.*

10. In experiments 1, 2, 3, use about 1 gm. of  $\text{CuSO}_4, 5\text{H}_2\text{O}$ , and about 8c.c. of water. In each case make two solutions exactly alike, except in the condition in question. 1. Does copper sulfate dissolve more quickly with or without agitation? 2. Does it dissolve more quickly in hot or cold water? 3. Does it dissolve more quickly powdered or in

lump? In 4 use carbonate of soda. 4. Does hot or cold water dissolve more of this salt? Dissolve all possible in 8c.c. of cold water (cold saturated solution), now add nearly as great a bulk of the carbonate as the bulk of the liquid; heat. 5. What happens as the solution (4) cools? (If nothing happens, it is because you failed to get a cold saturated solution.) 6. Into 8c.c. of water put about 1 gm. of plaster of Paris ( $\text{CaSO}_4$ ); you cannot see whether any has dissolved or not. Filter a little of the liquid upon a fragment of clean glass (not your cover glass) and gently heat till the water evaporates. Had any of the  $\text{CaSO}_4$  dissolved? 7. A *supersaturated* solution. Fill a test tube with small crystals of sodium thiosulfate ( $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$  — the photographer's "hypo"). Heat 2 or 3c.c. of water (about one-seventh the bulk of the "hypo"), pour it upon the crystals, dissolve them all by heating, bring to boiling, cork securely, and allow the liquid to become cold again. Then remove the cork; if nothing happens, drop in one very small crystal. Rapid crystallization will take place, accompanied by a temperature change. 8. With thermometer take the temperature of about 50c.c. of water in a beaker, put in a spoonful of saltpeter ( $\text{KNO}_3$  potassium nitrate) and take minimum temperature. 9. Into a beaker put about 50c.c. of water, boil upon the wire gauze and note the temperature when fully boiling. Save for 10. 10. Now put in about two teaspoonfuls of common salt ( $\text{NaCl}$  — sodium chlorid) and take the boiling temperature of this solution.

***Hydrogen Dioxid or Peroxid ( $\text{H}_2\text{O}_2$ ).***

11. Take about 10c.c. to your table. 1. Taste a few drops. 2. Put about 2 gm. of  $\text{MnO}_2$  in a test tube and add





5c.c. of  $\text{H}_2\text{O}_2$ ; prove that the escaping gas is oxygen. 3. To 2 or 3c.c. of solution of potassium permanganate ( $\text{KMnO}_4$ ) add a few drops of dilute  $\text{H}_2\text{SO}_4$  and then add 2 or 3c.c. of  $\text{H}_2\text{O}_2$ ; prove that O escapes. 4. Gently warm the remainder of the  $\text{H}_2\text{O}_2$  with a little freshly made lead sulfid ( $\text{PbS}$ ). It gives O to the  $\text{PbS}$ , forming white  $\text{PbSO}_4$  (lead sulfate). Does O escape?

Equations:  $\text{H}_2\text{O}_2 = \text{H}_2\text{O} + \text{O}$ . (3) The action is complex; the final products are  $\text{K}_2\text{SO}_4$ ,  $\text{MnSO}_4$ ,  $\text{H}_2\text{O}$ , O. (4) ?  $\text{H}_2\text{O}_2 + \text{PbS} = \text{PbSO}_4 +$  .

**(Fluorin F) Hydrofluoric Acid (HF).**

**12.** (Do not let the HF come in contact with the fingers.) Coat a piece of glass about 2"x3" with paraffin; scratch through the paraffin some design upon the glass; lay a piece of filter paper upon the glass and wet it with HF; after fifteen minutes or more clean and examine the glass. To clean the glass rinse with water, scrape with a knife, and finally gently heat and wipe with a cloth.

Equations:  $\text{SiO}_2 + ? \text{HF} = ? \text{H}_2\text{O} + \text{SiF}_4$  represents in part the action.

**\* Chlorin (Cl).**

**13.** 1. Make Cl by  $\text{MnO}_2$ , 5 or 6 gm. and concentrated HCl, 5c.c., well mixed and slightly warmed in a test tube. When the tube is half-full of Cl (recognized by its greenish color), heat the end of a narrow strip of copper foil and plunge it quickly into the Cl without letting the copper get wet. If bubbles interfere, stop heating and wait for them to subside. The copper glows and forms cupric chlorid. Notice the color it produces in the lamp flame. Get rid of

\* In paragraphs 13, 14, 15, the gases produced are irritating and hoods are desirable, though with care it is possible to get along without them.

the contents of the tube at once by rinsing into the drain-pipe of the sink. 2. Make a less offensive solution of Cl by using about one teaspoonful of bleaching powder, 50c.c. of water, and adding 5c.c. of dilute acid (acetic, hydrochloric, or sulfuric). 3. Try the bleaching action of this solution upon litmus paper, printer's and writing inks, and its effect upon offensive material ( $\text{H}_2\text{S}$  water may be used), and save a little for 16, 4.

Equation: (1) ?  $\text{HCl} + \text{MnO}_2 = ? \text{H}_2\text{O} + \text{MnCl}_2 +$  , representing final result.  $\text{Cu} + \text{Cl}_2 =$  . Bleaching and disinfecting are largely due to these two reactions:  $\text{H} + \text{Cl} =$  ? .  $\text{H}_2\text{O} + \text{Cl}_2 = 2\text{HCl} +$  .

### *Hydrochloric Acid (HCl).*

14. 1. Put in a side-necked test tube, in the order given (why?), 4c.c. of water, 6c.c. concentrated  $\text{H}_2\text{SO}_4$  and 3 or 4 gm. of coarse rock salt ( $\text{NaCl}$ ); hold moistened litmus paper at the end of the delivery tube. 2. Also breathe across the end of the tube and notice the fog which is formed whenever  $\text{HCl}$  gas comes in contact with vapor of water. 3. After the first rapid evolution of gas gently warm the s. n. tube, if necessary (do not heat to the boiling point of water in this or any similar case where a gas is generated in the presence of water (why?)); now fill a perfectly dry gas jar (mouth upwards or downwards?) with  $\text{HCl}$  gas. Remember that you can get at the weight of a gas from its formula. (See Part II, Sec. 9.) It is difficult to know when the jar is full, so you must send a steady current of gas to the bottom of the jar for at least 10 minutes, keeping the jar loosely covered; now cover securely, insert the jar in the water tank and slip it off from the cover. 4. Moisten the inside of your other gas jar with a few drops of strong  $\text{NH}_4\text{OH}$ , and







send HCl from the s. n. tube into this jar; heavy clouds of  $\text{NH}_4\text{Cl}$  (a solid). Why not collect HCl over water as you did O and H?

Equations: (1)  $\text{NaCl} + \text{H}_2\text{SO}_4 = \text{HNaSO}_4 + \text{HCl}$  .  $2\text{NaCl} + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{HCl}$  . (3)  $\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}$  .

### *Bromin (Br).*

**15.** Upon 1 crystal of potassium bromid (KBr) put 2 or 3 drops of concentrated  $\text{H}_2\text{SO}_4$ , and gently warm; reddish brown Br and colorless fuming HBr are evolved. 2. Nearly fill a test tube with water; add 1 drop of Br (this must be done with a dropping tube and in a hood), shake the tube to show the solubility of the Br. 3. Try the bleaching power of a little of this Br water upon litmus. 4. Transfer the Br water to a small beaker and pass  $\text{H}_2\text{S}$  (hydrogen sulfid) rapidly through it. The  $\text{H}_2\text{S}$  is made in the s. n. test tube by the action of quite dilute  $\text{H}_2\text{SO}_4$  upon 2 or 3 gm. of FeS (ferrous sulfid). Notice the change in color and clearness and the effect upon litmus (compare with 3). HBr is formed and S is set free.

Equations: (1) like 14, 1, but  $2\text{HBr} + \text{H}_2\text{SO}_4 = ?$   $\text{H}_2\text{O} + \text{Br}_2 + \text{SO}_2$ .  
(4) perhaps  $10\text{Br} + 2\text{H}_2\text{S} + 4\text{H}_2\text{O} = 10\text{HBr} + \text{H}_2\text{SO}_4 + \text{S}$ .

### *Iodin (I).*

**16.** 1. Upon one crystal of KI put 2 or 3 drops of concentrated  $\text{H}_2\text{SO}_4$  and gently warm; I and HI are evolved. What is the color of I vapor? 2. Put one small scale of I into each of four test tubes; then put into the tubes respectively 2c.c. of water, alcohol, carbon disulfid ( $\text{CS}_2$ ), water solution of KI. In which of the liquids does the I dissolve, and what are the colors? 3. To a beaker of dilute starch

paste (see Chemicals) add a drop or two of either the alcoholic or the KI solution of I; here we have free I. What is the effect on the starch? 4. Add to another lot of starch paste a drop or two of fresh KI solution. In this there should be no free I. What is the effect on the starch? Now add Cl water from **13**, 3, first a drop, then 10c.c. or more; the liberated I forms a blue compound with starch, which more Cl destroys. 5. Test a sample of baking powder for starch.

Equations: (1) like **15**, 1. (4)  $KI + Cl =$  + .

### *Sulfur (S).*

**17.** 1. Put a piece of broken roll sulfur, as large as half a pea, into 2c.c. of  $CS_2$  and shake for a few minutes. Does it perceptibly dissolve? Could a substance be soluble and yet not wholly dissolve? Why? 2. Pour the liquid (from 1) upon the dry bottom of an inverted gas jar, and let it evaporate spontaneously. Notice shape of crystals. 3. Nearly fill a 2 oz. hessian crucible with broken roll sulfur and heat until all has melted. Do not allow it to burn at the top or sides of the crucible. Let it cool uncovered. As soon as the surface is just crusted over (watch it carefully), make a hole and quickly drain out all the liquid. Cut out the crust. Are the crystals like those in 2? 4. Put some fragments of roll sulfur into a large test tube (not side-necked), and apply heat, noting carefully changes in color and thickness which take place as the temperature rises. 5. When the sulfur is boiling, pour part of it into a beaker of cold water, and examine. Recall **5**, 2 and **6**, 2. 6. Ignite a little S. Does it take fire easily? Describe the flame. Where is the product of the combustion? Can you smell it? 7. Hold a strip of moistened blue litmus paper in





the fumes of the burning S. 8. Make a simple blow-pipe by bending about 8 in. of  $\frac{3}{16}$  glass tubing at a right angle 1 in. from the end and closing the shorter arm to about the caliber of a pin. Upon a piece of charcoal put a mixture of one part finely powdered FeS and four parts of dry  $\text{Na}_2\text{CO}_3$ ; the whole mixture should not measure more than  $\frac{1}{2}$  c.c.; fuse thoroughly with the blow-pipe flame. Put the fused mass upon a silver coin and moisten with water. Sodium sulfid is formed by the fusion, and this with the silver makes black silver sulfid. Give the history of the sulfur in this experiment.

Equations: (6)  $\text{S} + \text{O}_2 =$

(7)  $\text{SO}_2 + \text{H}_2\text{O} =$

### *Hydrogen Sulfid ( $\text{H}_2\text{S}$ ).*

**18.** Have everything ready for the whole paragraph before you generate  $\text{H}_2\text{S}$ . 1. Prepare  $\text{H}_2\text{S}$  as in **15**, 4. 2. Notice odor. 3. Direct the gas upon moistened litmus paper, red and blue, until you notice some change. 4. Direct the gas upon paper which has been wet in solution of acetate or nitrate of lead; the black product is PbS. 5. Collect a little  $\text{H}_2\text{S}$  over water—a test tube full—apply a light. It burns, with almost invisible flame, depositing some sulfur. 6. Mix in a test tube solutions of antimony chlorid ( $\text{SbCl}_3$ ) and ferrous chlorid ( $\text{FeCl}_2$ ). Make them distinctly acid with dilute HCl, if they are not already, and pass  $\text{H}_2\text{S}$  through; orange antimony sulfid ( $\text{Sb}_2\text{S}_3$ ) falls as an insoluble compound, or precipitate, but no FeS falls. Now filter and add  $\text{NH}_4\text{OH}$  to the filtrate till alkaline after stirring; in the presence of the  $\text{NH}_4\text{OH}$  black FeS is formed and precipitated. All the metals that act like antimony can be separated from all the metals that act like iron in this way; and this and similar methods are largely used in the application of chemistry, called Qualitative Analysis.

Equations: (1)  $\text{FeS} + \text{H}_2\text{SO}_4 = \quad + \quad$ . (4)  $\text{H}_2\text{S} + \text{Pb}(\text{NO}_3)_2 =$   
 $? \text{HNO}_3 + \quad$ . (5)  $\text{H}_2\text{S} + ? \text{O} = \text{H}_2\text{O} + \text{SO}_2$  and  $\text{H}_2\text{S} + \text{O} = \text{S} + \quad$ .  
 (6)  $? \text{SbCl}_3 + ? \text{H}_2\text{S} = ? \text{Sb}_2\text{S}_3 + ? \text{HCl}$ .  $\text{NH}_4\text{OH} + \text{HCl} = \text{NH}_4\text{Cl} + \quad$ .  
 $? \text{NH}_4\text{OH} + \text{H}_2\text{S} = (\text{NH}_4)_2\text{S} + \quad$ .  $(\text{NH}_4)_2\text{S} + \text{FeCl}_2 = \quad + \quad$ .

### *Sulfurous Oxid ( $\text{SO}_2$ ).*

**19.** See **17**, 6. Before generating  $\text{SO}_2$  have all the materials for the paragraph ready. 1. Put in a side-necked test tube 1 sq. in. of sheet copper, not more than 10cc. of concentrated  $\text{H}_2\text{SO}_4$  and apply heat until the gas begins to come, then regulate the heat carefully. On no account dip the delivery tube into any liquid, as the sucking back of water upon the hot  $\text{H}_2\text{SO}_4$  would produce a dangerous explosion. Get information as to odor, color, relative weight, reaction, and combustibility of  $\text{SO}_2$ . 2. Put in a gas jar a few c.c. of logwood solution, let the  $\text{SO}_2$  run into the jar (mouth upwards), then cover and shake. [3. If obtainable, a blue or purple flower may be put in the jar instead of logwood to show bleaching effect.] 4. Put in another jar a dilute solution of  $\text{KMnO}_4$  and let the  $\text{SO}_2$  run into this jar as in 2. The  $\text{SO}_2$  takes oxygen, finally forming sulfates of potassium and manganese. Save part of the liquid in the jar for **20**, 7. 5. Let the contents of the s. n. test tube cool, pour out the liquid and remove any copper that is left; what do you find at the bottom of the tube? Why is it not blue? What is the effect on the color of adding water? 6. What gas is evolved by the action of dilute  $\text{H}_2\text{SO}_4$  upon a little solid sulfite of sodium ( $\text{Na}_2\text{SO}_3$ ) in a test tube?

Equations: (1) final result  $\text{Cu} + ? \text{H}_2\text{SO}_4 = \text{CuSO}_4 + \quad + ? \text{H}_2\text{O}$ .  
 (4) perhaps  $3\text{SO}_2 + 2\text{KMnO}_4 = \text{K}_2\text{SO}_4 + 2\text{MnSO}_4 + \text{O}_2$ . (6)  $\text{Na}_2\text{SO}_3 +$   
 $\text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + \quad + \quad$ .







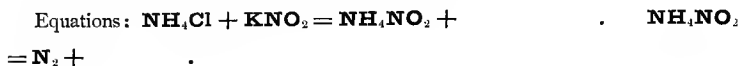
**Sulfuric Acid ( $H_2SO_4$ ).**

**20.** 1. Put about 2c.c. of water in a test tube, and add an equal volume of concentrated  $H_2SO_4$ , holding the lower part of the tube in the hand. (Remember in diluting concentrated  $H_2SO_4$  always to add the acid to the water, and slowly.) 2. Make a very thick syrup of sugar and hot water (test tube), and after it has nearly cooled add an equal volume of concentrated  $H_2SO_4$ , holding the tube by a test tube holder. 3. Trace some design upon white paper with dilute  $H_2SO_4$ , and gently warm the paper until it is dry. 4. Add dilute HCl and then baric chlorid solution ( $BaCl_2$ ) to dilute  $H_2SO_4$ . 5. Repeat 4, using a solution of  $Na_2SO_4$  instead of  $H_2SO_4$ . The precipitate in 4 and 5 is barium sulfate ( $BaSO_4$ ), and this is a common method of identifying sulfates. The HCl has no effect here, but without the HCl a number of other substances — carbonates, phosphates etc. — might be mistaken for sulfates. 6. Try phosphate of soda solution and  $BaCl_2$ , with and without HCl. 7. Test the liquid saved from **19**, 4, for sulfates. [8. Repeat **15**, 4, and ascertain whether any sulfuric acid is formed.] Very few illustrations of the properties of this important substance are given here because illustrations are all through the book. How do experiments **7**, 1, **8**, 6, **14**, 1, **15**, 1, **16**, 1, **18**, 1, **19**, 1, 4, 6, illustrate properties of  $H_2SO_4$ ? After you have done the following experiments, note here the properties of  $H_2SO_4$  which they illustrate: **23**, 5, **28**, 1, 2, **36**, 6, **38**, 1, **39**, 1, **47**.

Equations: (1) perhaps  $H_2SO_4 + 2H_2O = H_6SO_6$ . (2)  $C_{12}H_{22}O_{11}$  is partly broken into carbon and water. (4)  $H_2SO_4 + BaCl_2 + nHCl = nHCl + BaSO_4 +$  . (5)  $Na_2SO_4 + BaCl_2 + nHCl =$   
 $+ +$  .

*Nitrogen (N).*

**21.** Put in a s. n. test tube 2 gm. of  $\text{NH}_4\text{Cl}$ , 3 gm. of potassium nitrite ( $\text{KNO}_2$ ), just moisten with a few drops of water. Apply gentle heat and collect a jar of N over water. Does N maintain combustion or burn?

*Air.*

**22.** Take about 8 in. of one-fourth or five-eighths in. glass tubing; close one end and bend about one inch to a right

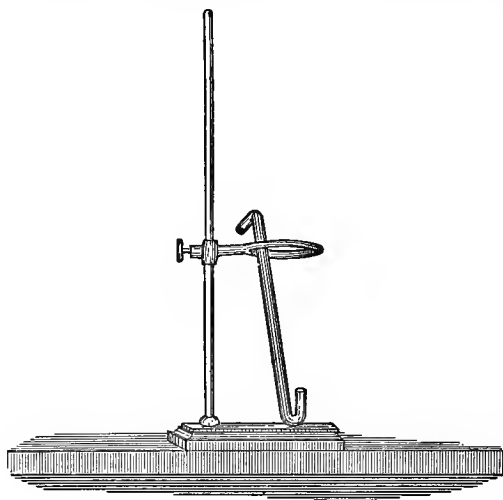


Fig. 6.

angle; bend the open end in a short loop (Fig. 6). After the tube is cool slip to the closed end a piece of phosphorus (P) as large as a grain of rice. (Dry the P with filter paper, do not rub or squeeze it, and do not at any time touch it with the fingers. Why?) Fill the open loop with water, mark

its height with a file, then stand the tube upright in a beaker of water, the open end half an inch or so under water. Now heat the P gently for several minutes, and





let the whole stand until it is cold again. (Why?) Observe that water has filled almost exactly one-fifth of the air space in the tube. The P has exhausted the oxygen.

2. Demonstrate the presence of carbon dioxide ( $\text{CO}_2$ ) in the air by exposing half an hour or more 100 c.c. of lime water ( $\text{Ca}(\text{OH})_2$ ) in an open beaker;  $\text{CaCO}_3$  is formed on the surface.

3. Half-fill two test tubes with beef tea. Heat both to boiling, then plug the mouth of one loosely with a wad of cotton that has just been passed through the lamp flame two or three times. Germs from the atmosphere will cause putrefaction and mould in the open tube but not in the plugged tube. In the laboratory this process is likely to be very slow — two or three weeks.

Equations: (1)  $\text{P}_2 + ? \text{O} = \text{P}_2\text{O}_5$ .  $\text{P}_2\text{O}_5 + \text{H}_2\text{O} = ? \text{HPO}_3$ , or,  
 $\text{P}_2\text{O}_5 + ? \text{H}_2\text{O} = 2\text{H}_3\text{PO}_4$ . (2)  $\text{CaO}_2\text{H}_2 + \text{CO}_2 = \quad + \quad$ .

### *Ammonia* ( $\text{NH}_3$ ).

**23.** 1. Grind together in a mortar about 2 gm. of  $\text{NH}_4\text{Cl}$  and 4 gm. of  $\text{CaO}$ ;  $\text{NH}_3$  is set free, to be recognized by its odor, its effect upon moist red litmus paper and the fumes when a glass rod moistened with concentrated  $\text{HCl}$  is brought near. Save for 3. 2. Put in a test tube a solution of any ammonium salt ( $\text{NH}_4\text{Cl}$  or  $(\text{NH}_4)_2\text{SO}_4$ ), add a few drops of  $\text{NaOH}$  solution and warm; is  $\text{NH}_3$  liberated? How can you tell  $\text{NH}_3$  from any other gas? 3. Transfer the mixture from 1 to a s. n. test tube; just moisten it and collect, with heat, a dry jar of  $\text{NH}_3$  and show its absorption by water. The experiment is exactly like the last part of **14**, 3, except, which way should the mouth of the jar be — up or down — to collect a gas like  $\text{NH}_3$ ? [4. Get a s. n. test tube all ready for the preparation of  $\text{O}$  (**8**, 3); put in a beaker about 200 c.c.

of strong ammonium hydroxid ( $\text{NH}_4\text{OH}$ ); set it upon the sheet iron or gauze and gently heat till bubbles of ammonia come rapidly

away, then make O not too fast and have it pass through the liquid in the beaker; when both gases are coming well, apply a light to the mouth of the beaker. If you have managed this well, you will have a brilliant combustion of O &  $\text{NH}_3$ .] 5. *Neutralization*. Recall 14, 4. To 100 c.c. of dilute  $\text{H}_2\text{SO}_4$  add dilute

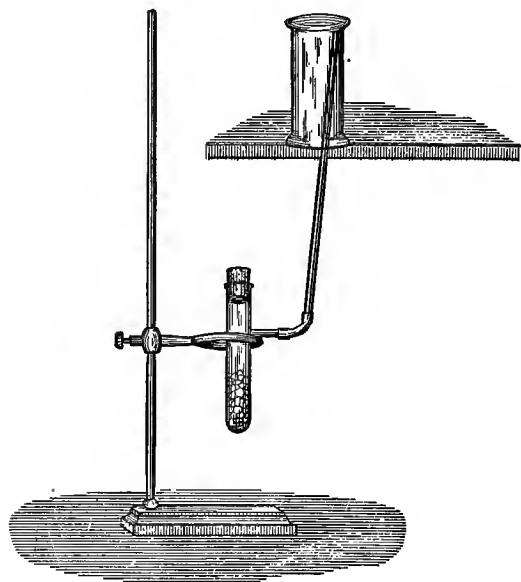
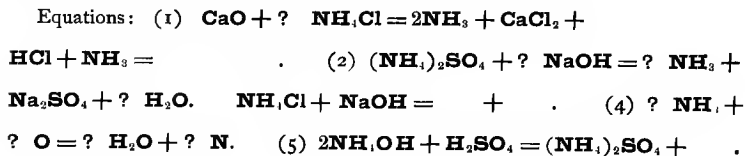


Fig. 7.

$\text{NH}_4\text{OH}$  until neither red nor blue litmus is affected by the liquid. Use only one piece of litmus paper until the end of the operation. Try not to use too much  $\text{NH}_4\text{OH}$ ; but if you do, a few drops of the acid may be added, etc. Save for 24, 2.









**Compounds of Ammonium ( $\text{NH}_4$ ).**

**24.** 1. To 1c.c. of solutions of chlorid of cadmium and chlorid of tin, in separate test tubes, add first 3 or 4 drops, and then 10c.c. of ammonium sulfid ( $(\text{NH}_4)_2\text{S}$ ); warm; in both cases a sulfid is formed, but the tin sulfid dissolves in the excess and the cadmium sulfid does not. A similar reaction is used in Qualitative analysis to separate several metals that act like tin from several that act like cadmium.

2. Evaporate the neutral solution of  $(\text{NH}_4)_2\text{SO}_4$  (**23**, 5) nearly to dryness, let it cool; small crystals of  $(\text{NH}_4)_2\text{SO}_4$  will form. See Part IV, 9. In a manner similar to this, most ammonium salts can be made. How could you make  $\text{NH}_4\text{NO}_3$ ? Examine the various ammonium salts at hand; do they smell of  $\text{NH}_3$ ?

**Nitrogen Monoxid ( $\text{N}_2\text{O}$ ).**

**25.** Collect over water one jar of  $\text{N}_2\text{O}$ , made by heating about 4 gm. of  $\text{NH}_4\text{NO}_3$  in a s. n. test tube. The  $\text{NH}_4\text{NO}_3$  melts to a liquid before the gas comes; after the evolution of gas begins, use just enough heat to keep up a steady current. Refer again to cautions in **8**, 3. Observe sweetish taste and odor and effect upon combustion (stick).

Equations:  $\text{NH}_4\text{NO}_3 = \quad + \quad$ . Compare with the decomposition of  $\text{NH}_4\text{NO}_2$  in **21**. In combustion  $\text{N}_2\text{O} = \text{N}_2 + \text{O}$ .

**Nitrogen Dioxid ( $\text{NO}$ ).**

**26.** 1. Equip a 250c.c. Erlenmeyer flask with delivery tube and funnel tube, as in Fig. 8. See that all connections are tight. Slide into the flask (do not drop it in) about 2 sq. in. (6 or 7 gm.) of sheet copper; now add through the funnel tube 20–30 c.c. of  $\text{HNO}_3$  sp. gr. 1.2. The action may

be started by warming, but do not continue the heat. Collect two jars of the gas; one is to be used in 2 and must be full; the other is to be used in 27 and need not be full, and is not

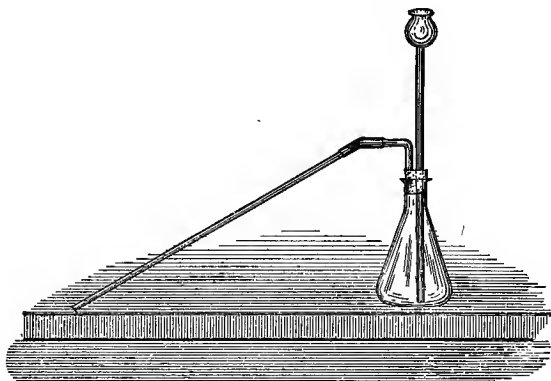


Fig. 8.

to be taken from the water tank. 2. Remove the full jar from the tank; slide the cover off the least possible degree and pour in 1 c.c. (and not a drop more) of  $\text{CS}_2$ . Shake thoroughly, uncover completely, and at the same instant apply a light. (This "flashlight" affects a photographic plate in a dark room.)

Equations: (1)  $3\text{Cu} + 8\text{HNO}_3 = ? \quad \text{Cu}(\text{NO}_3)_2 + ? \quad \text{NO} + ? \quad \text{H}_2\text{O}$ .

(2) Something like this:  $\text{CS}_2 + 4\text{NO} = \text{CO}_2 + \text{SO}_2 + 4\text{N} + \text{S}$ .

### *Nitrogen Tetroxid ( $\text{NO}_2$ ).*

27. Arrange for the preparation of O (8, 3). Send the O very slowly and intermittently into the jar of NO from 26. Notice the immediate formation of the brown gas ( $\text{NO}_2$ ) and its rapid absorption by water. If you manage





this skilfully, you can carry the water nearly to the top of the jar.



### *Nitric Acid ( $\text{HNO}_3$ ).*

**28.** [1. Where the conditions are suitable a little strong  $\text{HNO}_3$  may be made by putting 20 gm. of common saltpeter ( $\text{KNO}_3$ ) and 20 gm. of concentrated  $\text{H}_2\text{SO}_4$  into a 2 to 4 oz. glass-stoppered retort (Fig.

9) and gently distilling in the hood. This must not be done in the open laboratory or with corks and rubber connections.] 2. Into a test tube put 2c.c. solution of  $\text{KNO}_3$ , 1c.c. concentrated  $\text{H}_2\text{SO}_4$ , and a strip of copper foil. This makes and uses up  $\text{HNO}_3$  in the same operation. Nitrates may generally be recognized by this

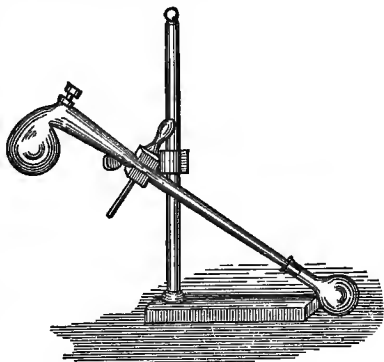


Fig. 9.

reaction. 3. Into 2c.c. of concentrated  $\text{HNO}_3$  put a scrap of tin; the tin is oxidized to a white, nearly insoluble powder ( $\text{H}_2\text{SnO}_3$ , metastannic acid), while in 2 and **26**, 1 the copper enters into the formation of a soluble nitrate. 4. Boil a little finely powdered  $\text{FeS}$  in about 2c.c. of concentrated  $\text{HNO}_3$  carefully for two or three minutes, dilute with water to 10 or 15c.c., filter, and add solution of  $\text{BaCl}_2$  or  $\text{Ba}(\text{NO}_3)_2$ . A white precipitate. What does it signify? See **20**, 4 and 5. [4. Put in test tube about 4c.c. concentrated  $\text{HNO}_3$ ; loosely plug the mouth of the tube with horse-

hair, bring the  $\text{HNO}_3$  to boiling till its vapor comes in contact with the horsehair. Hood only.]

Equations: (1)  $2\text{KNO}_3 + \text{H}_2\text{SO}_4 = ? \text{HNO}_3 +$  . With less  
heat and nitrate,  $\text{KNO}_3 + \text{H}_2\text{SO}_4 = \text{H}_2\text{KSO}_4 +$  . (2) like 1 and  
26, 1.

### *Phosphorus (P).*

**29.** (See cautions in **22**, 1. Do not let it become dry until you wish to ignite it. Have water at hand to extinguish it if it inflames prematurely. It can safely be cut or handled under water). 1. Take a piece of yellow P the size of a grain of rice, partially dry it; notice fumes and odor. 2. Put it into 4 or 5 c.c. of water and warm. Notice its low melting point—about  $44^\circ\text{C}$ . 3. Pour it into cold water, cut it in two, dry half, dissolve it in 2 c.c. of  $\text{CS}_2$  (dry test tube). This is its best solvent. 4. Pour this solution upon a strip of filter paper, letting the drops fall into the sink and leaving enough of the paper dry to hold it by; wave the paper back and forth; the  $\text{CS}_2$  evaporates and the finely divided P ignites spontaneously. Notice that generally the paper does not take fire, which illustrates the low kindling point of P. 5. Dry the other half of the P, put it on a crucible cover, ignite it with a warm glass rod and cover at once with a perfectly dry beaker. The success of this depends upon the dryness of everything. See the white, flaky product— $\text{P}_2\text{O}_5$ . 6. Notice also how deliquescent this oxid is; breathe on it. 7. Wash the inside of the beaker with a few c.c. of water. Is this solution strongly acid to litmus? It contains  $\text{HPO}_3$  metaphosphoric acid. [8. Does “red” or “amorphous” P melt, dissolve, or ignite as the other form did?]

Equations:  $2\text{P} + ? \text{O} = \text{P}_2\text{O}_5$ .  $\text{P}_2\text{O}_5 + \text{H}_2\text{O} = ? \text{HPO}_3$ .







***Hydrogen Phosphid ( $H_3P$ ).***

[30. This gas may be made by a number of methods, of which the simplest is : Drop into dilute  $HCl$  in a beaker or conical test glass pieces of calcium phosphid ( $Ca_3P_2$ ) of the size of a pea until it comes well. Notice odor and spontaneous inflammability. (Hood only.) Let the pupil find other methods in the text-books.

Equation :  $Ca_3P_2 + 6HCl = ? CaCl_2 + \quad .]$

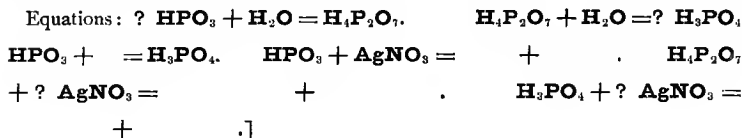
***Hypophosphorous Acid ( $H_3PO_2$ ).***

31. 1. To 3c.c. of solution of  $CuSO_4$  add 2c.c. dilute  $H_3PO_2$  and gently warm until no more change occurs. During the change are formed cuprous hydrid ( $Cu_2H_2$ ), cupric hydrid ( $CuH_2$ ), copper ( $Cu$ ), phosphorous acid ( $H_3PO_3$ ). It is a powerful reducing agent. See 7, 8. 2. Try in the same manner the reducing action of  $H_3PO_2$  upon  $AgNO_3$  solution.  $AgH$  is produced which, upon charcoal with the blow-pipe, will give a little globule of silver.

***The Phosphoric Acids ( $HPO_3$ ,  $H_4P_2O_7$ ,  $H_3PO_4$ ).***

[32. 1. Put about 2 gm. of "glacial" or metaphosphoric acid ( $HPO_3$ ) into an evaporating dish, add 2 or 3c.c. of water, and after a little of the  $HPO_3$  has dissolved pour the liquid (only) into a test tube, now add  $AgNO_3$  solution; white precipitate of  $AgPO_3$ . 2. Now to the  $HPO_3$  left in the dish add 20 to 30c.c. of water, dissolve with heat and boil a minute or two after complete solution. Take a little of the liquid out and add  $AgNO_3$ ; little or no precipitate until you add a drop or two of  $NH_4OH$ , when white  $Ag_4P_2O_7$  appears. 3. Now fill the dish with water, keep it full, and boil half an hour or more : test a portion of this solution with  $AgNO_3$ ,

adding the merest drop of  $\text{NH}_4\text{OH}$  and you will get light yellow  $\text{Ag}_3\text{PO}_4$ . These experiments mark the progress from  $\text{HPO}_3$  to  $\text{H}_3\text{PO}_4$  by combination with water. All of the precipitates are soluble in  $\text{NH}_4\text{OH}$ , so you must be careful not to add much.



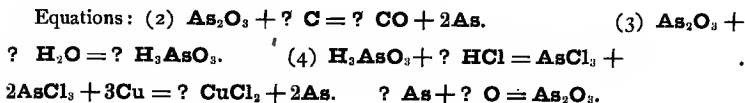
### *Arsenic (As).*

**33.** Examine the element arsenic and such compounds as are on the shelf. 1. Draw out 8 in. of  $\frac{1}{4}$  in. glass tubing so as to make two 4 in. tubes each closed at one end. Convey to the closed end of one of the tubes about  $\frac{1}{10}$  gm. of white arsenic (arsenious oxid,  $\text{As}_2\text{O}_3$ )—about what you could put on the tip of a penknife. Now heat and notice that the  $\text{As}_2\text{O}_3$  becomes a vapor and condenses in glittering (crystalline) form on the cooler portion of the tube (it *sublimes*). Do not heat long enough to drive any vapor out of the tube, as it is very poisonous. Save tube for 3. 2. Intimately mix the same amount of  $\text{As}_2\text{O}_3$  with powdered charcoal, put it in the other tube, put about an inch of charcoal above it, then heat first the charcoal, then the mixture. You will need the heat of a Bunsen burner here. After sufficient heat you will find on the tube a dark shining band or “mirror” of elemental arsenic, reduced by the charcoal. 3. Cut off the portion of the tube (1) containing the  $\text{As}_2\text{O}_3$  and boil it in 10 to 20 c.c. of water in the evaporating dish; test its reaction (see 9, 9). 4. Now add a few drops of  $\text{HCl}$  and a slip of bright copper foil and boil again; the arsenic is deposited on the copper, *alloying*



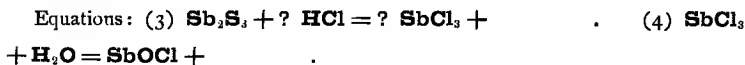


with it. 5. Dry the copper thoroughly with filter paper and heat it in a perfectly clean and bright test tube; a faint ring of  $\text{As}_2\text{O}_3$  crystals appears just above the hot portion of the tube. (See what you can find about the Marsh-Berzelius test for arsenic).



### *Antimony (Sb).*

**34.** 1. Upon a piece of Sb as large as a grain of rice try the effect of hot concentrated HCl. 2. After rinsing try upon the same piece the effect of concentrated  $\text{HNO}_3$ . Product is  $\text{HSbO}_3$ . Compare 28, 3. 3. Dissolve with heat 1 gm. of  $\text{Sb}_2\text{S}_3$  in 3 or 4c.c. concentrated HCl; what gas escapes? Evaporate the solution nearly to dryness; the product is  $\text{SbCl}_3$ . 4. Pour this thick liquid into a beaker of water; white  $\text{SbOCl}$  is produced. Occasionally, when the evaporation is insufficient, an orange instead of a white precipitate is obtained in 4. What is it? See 18, 6.



### *Boric Acid ( $\text{H}_3\text{BO}_3$ ).*

**35.** 1. Make a loop in a platinum wire by winding one end of it halfway up the tip of a lead-pencil. Melt into this loop little bits of borax until the loop is just filled with a clear glassy substance. Now melt thoroughly into this glass with your blow-pipe, pieces no larger than the head of a pin, of salts of either iron, copper, cobalt, nickel, manganese, chromium. Each of these gives a characteristic color. If

you have not the platinum wires, you may use an iron wire, but the bead will then be colored at the start from the iron. Look up "Borax Bead." 2. Using about 10c.c. of water, make a *hot saturated* (?) solution of borax ( $\text{Na}_2\text{B}_4\text{O}_7$  is "dry" borax) in evaporating dish; transfer it to a beaker and add concentrated  $\text{HCl}$  *in excess* (?).  $\text{H}_3\text{BO}_3$  separates in pearly scalelike crystals; when the mixture is cold transfer it to a filter and wash three or four times with a very little cold water. 3. Try the solubility of some of these crystals in boiling water. 4. Put 10c.c. of alcohol in an evaporating dish; add some of the  $\text{H}_3\text{BO}_3$  crystals, ignite the alcohol, stir and notice the green flame characteristic of  $\text{H}_3\text{BO}_3$ . (The flame may be tinged with yellow from the presence of the sodium.)

Equation: (2)  $\text{Na}_2\text{B}_4\text{O}_7 + ? \text{HCl} + ? \text{H}_2\text{O} = ? \text{NaCl} + ? \text{H}_3\text{BO}_3$ .

### ***Compounds of Silicon (Si).***

**36.** 1. To 1c.c. of solution sodium silicate ( $\text{Na}_4\text{SiO}_4$ ) add an equal amount of concentrated  $\text{HCl}$ ; gelatinous  $\text{H}_4\text{SiO}_4$  (silicic acid) forms at once. 2. Put in a beaker about 16c.c. of water, add just half as much  $\text{Na}_4\text{SiO}_4$  solution, and acidulate with dilute  $\text{HCl}$ .  $\text{H}_4\text{SiO}_4$  remains in solution. (What is the other product?) 3. Transfer the mixture (2) to an evaporating dish and bring to dryness. At a certain stage the liquid will become gelatinous; at this point the evaporation must be carried on with *constant stirring* and *very gentle* heat, or everything will jump out of the dish. The end is reached when a dry white powder is left in the dish. After the dish has cooled, add water, filter, wash the precipitate and dry it in the paper. Gritty, white  $\text{SiO}_2$  is obtained. What ran through in the filtrate? Prove by tasting. 4. Try the solubility of this  $\text{SiO}_2$  in boiling  $\text{NaOH}$  solution.







5. Acidulate liquid from 4 with HCl. It contains  $\text{Na}_4\text{SiO}_4$  and will act like 1 or 2 according to its strength. Why do the stoppers of NaOH bottles stick? [6. Take pure sand, or powdered glass, or  $\text{SiO}_2$  made as in 2 & 3, about 2 gm., fluor-spar 2 gm., mix, put in a test tube, add enough concentrated  $\text{H}_2\text{SO}_4$  to make a paste, adjust cork and delivery tube; put in another test tube  $\frac{1}{2}$  in. of mercury; dip the delivery tube into the mercury and keep it there; nearly fill the second tube with water, finally gently heat the first tube.  $\text{SiF}_4$  is evolved; as it passes into the water it forms hydro-fluor-silicic acid ( $\text{H}_2\text{SiF}_6$ ) and gelatinous  $\text{H}_4\text{SiO}_4$ . Save the mercury. 7. Filter some of the liquid (6), and notice that the filtrate is strongly acid. Add to it a solution of some potassium salt, e. g. KCl; precipitate of  $\text{K}_2\text{SiF}_6$  (not in very dilute solutions).]

Equations: (1)  $\text{Na}_4\text{SiO}_4 + ? \text{HCl} = \text{H}_4\text{SiO}_4 + \quad .$  (3)  $\text{H}_4\text{SiO}_4 - \text{H}_2\text{O} = \quad .$   $\text{H}_4\text{SiO}_4 - 2\text{H}_2\text{O} = \quad .$  (4)  $\text{SiO}_2 + ? \text{NaOH} = \text{Na}_4\text{SiO}_4 + \quad .$  (6)  $\text{CaF}_2 + \text{H}_2\text{SO}_4 = \quad + \quad .$   $\text{SiO}_2 + ? \text{HF} = \quad + ? \text{H}_2\text{O}.$   $? \text{SiF}_4 + ? \text{H}_2\text{O} = \text{H}_4\text{SiO}_4 + ? \text{H}_2\text{SiF}_6.$  (7)  $\text{H}_2\text{SiF}_6 + ? \text{KCl} = ? \text{HCl} + \quad .$

### Carbon (C).

**37.** Examine the various forms of carbon. 1. Put about 2 teaspoonfuls of powdered bone black into a filter, making the center concave; filter slowly through this 20 or 30 c.c. of logwood solution. 2. Upon a piece of charcoal put  $\frac{1}{2}$  gm. or less of PbO and direct with the blow-pipe the inner lamp flame upon it. Illustrates reducing action of C. See **33**, 2; also compare with **7**, 8, and **31**, 1, 2.

Equation: (2)  $\text{PbO} + \text{C} = \text{CO} + \text{Pb}.$   $\text{CO} + \text{O} = \quad .$

*Carbon Monoxid (CO).*

**38.** 1. Put in a s. n. test tube 2c.c. of formic acid ( $\text{H}_2\text{CO}_2$ ) and 4 or 5c.c. of concentrated  $\text{H}_2\text{SO}_4$ ; gently warm. Collect one jar of the gas over water for 2. Pass the gas through lime water—about 20c.c. in a beaker; no result. 2. Removing the jar (1) covered, uncover and apply a light. Where have you seen this blue flame before? 3. Pour the lime water (1) into the jar and shake; it turns milky now, showing the presence of  $\text{CO}_2$ .

Equations: (1)  $\text{H}_2\text{CO}_2 + \text{H}_2\text{SO}_4 = (\text{H}_2\text{SO}_4 + \text{H}_2\text{O}) + \text{CO}$ . (2)  $\text{CO} + \text{O} =$  . For 3 see next paragraph.

*Carbon Dioxid (CO<sub>2</sub>).*

**39.** 1. Put a small crystal of  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  into each of two test tubes; to one add dilute  $\text{HCl}$ , to the other dilute  $\text{H}_2\text{SO}_4$ ; rapid escape of  $\text{CO}_2$ . What are the other products? 2. Put in the Erlenmeyer flask (26, 1) about 10 gm. of marble dust ( $\text{CaCO}_3$ ); cover with water and add concentrated  $\text{HCl}$  (10 to 20 c.c.), until the gas comes well. Put about 20c.c. of lime water ( $\text{Ca}(\text{OH})_2$ ) in a beaker, test its reaction, and make the  $\text{CO}_2$  pass through it until it becomes first milky and finally clear or almost clear again. Try the reaction when it is clear. There is produced first insoluble  $\text{CaCO}_3$  and later soluble  $\text{H}_2\text{Ca}(\text{CO}_3)_2$ . Save for 3. Collect two jars of  $\text{CO}_2$  over water. Remove one, covered, turn the mouth the right way (?), uncover and plunge into the jar a splinter which has been dipped in turpentine and lighted. 4. Hold a lighted match near the surface of the table and pour the gas in the other jar rapidly upon it. 5. Put the beaker (2) on the gauze and boil the liquid. As the excess of  $\text{CO}_2$  is boiled out,  $\text{CaCO}_3$  is again deposited, though in a





denser form than in 2. 6. Hold for an instant a beaker inverted above the lamp flame, then test the air in it for  $\text{CO}_2$  (with a little  $\text{Ca}(\text{OH})_2$ ). 7. Lower a little piece of burning wood into a gas jar and test the air in it for  $\text{CO}_2$ . 8. Prove that the air that comes from the lungs contains  $\text{CO}_2$ .

Equations: (1)  $\text{Na}_2\text{CO}_3 + \text{HCl} = \text{HNaCO}_3 +$  , in very dilute solutions.  $\text{Na}_2\text{CO}_3 + 2\text{HCl} = \text{H}_2\text{O} +$  + .  $\text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 +$  + . (2)  $\text{CaO}_2\text{H}_2 + \text{CO}_2 = \text{CaCO}_3 +$  .  $\text{CaCO}_3 + \text{CO}_2 +$  =  $\text{H}_2\text{Ca}(\text{CO}_3)_2$ . (5)  $\text{H}_2\text{Ca}(\text{CO}_3)_2 = \text{CaCO}_3 + \text{H}_2\text{O} +$  .

### *Carbon Disulphid ( $\text{CS}_2$ ).*

**40.** Recall its use as a solvent in **16**, **2**, **17**, **1**, **29**, **3**; also see **26**, **2**. 1. In 1 c.c. of  $\text{CS}_2$  shake for a minute or two a little chip of caoutchouc; pour the liquid upon the bottom of an inverted gas jar or beaker and let it evaporate. A film of rubber is left. 2. Into a dry jar put 1 c.c. (not any more) of  $\text{CS}_2$ ; cover, shake, and plunge into the jar a hot glass rod. This shows the low kindling point of  $\text{CS}_2$ . What gas can you smell in the jar? [3. Take about 6 in. of  $\frac{1}{2}$  in. glass tubing, close one end, and bend near the closed end like Fig. 10. When it is cool pour to the closed end 1 c.c. of  $\text{CS}_2$ . Keep horizontal and convey about 1 gm. of zinc dust to the middle of the tube. Now heat the Zn very hot. Keep it hot, then cautiously wave the flame under the  $\text{CS}_2$  till its

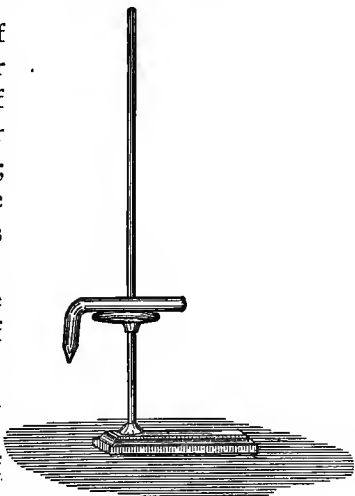


Fig. 10.

vapor passes over the Zn. Do not put much heat on the  $\text{CS}_2$  or it will suddenly jump out of the tube.  $\text{ZnS}$  is formed with a glow and C deposits on the tube. Can you get  $\text{H}_2\text{S}$  from the  $\text{ZnS}$  and  $\text{HCl}$ ? See 18, 4.]

Equations: (2)  $\text{CS}_2 + ? \text{O} = \text{CO}_2 + \text{SO}_2 + \text{S}$ , perhaps. (3)  $\text{CS}_2 + 2\text{Zn} = ? \text{ZnS} + \text{ZnS} + 2\text{HCl} = +$ .

### *The Lamp Flame.*

41. 1. Examine the lamp or candle flame and identify three somewhat distinct cones. 2. Holding a piece of white paper horizontally, carry it down upon the flame nearly to the wick and bring it up and away before it can take fire. The paper is charred in a circle. 3. Hold a thin stick across the flame and remove it before it burns. It is charred in two places, the part within the flame being affected not at all, or very little. 4. Quickly carry the phosphorus end of a match through the flame to the center of the wick, and with great quickness remove it. With skill enough the match can be brought out without igniting. 5. Take 2 in. of  $\frac{1}{4}$  in. glass tubing, hold it by tongs with the open end in the center of the flame at an angle of about  $45^\circ$ ; see if you can light the vapor that comes through the tube. From these experiments what do you learn about the lamp flame? 6. Examine the Bunsen burner, and try it, getting the luminous smoky flame and the blue non-luminous flame. 7. While you have the Bunsen burner (luminous flame), carry some wire gauze halfway down the flame; does the flame pass through? 8. Try blowing out the gas and lighting it between two pieces of gauze held horizontally. 9. Do 7 and 8 work as well with the luminous flame? 10. Put a bit of antimony upon a clean fragment of porcelain and laying it







upon charcoal or holding it with the test tube holders direct upon the antimony the extreme tip of the blow-pipe flame. More or less of the yellow-white  $\text{Sb}_2\text{O}_3$  is produced. 11. Now bring the porcelain a little nearer and use the inner flame. The oxygen is removed in patches, leaving a coating of antimony, dark and shining. These experiments illustrate oxidizing and reducing flames. 12. Blow from a dry glass tube, at a distance of a few inches, about 1 c.c. of lycopodium upon the lamp flame. Dust explosion.

Paragraphs 42-48 have reference to the formation and some of the properties of the more familiar metallic compounds. In all cases the bearing of the references upon the subject of the paragraph should be written in the notes, though the experiments referred to need not be repeated.

### *Oxids (Compounds with O).*

**42. Formation:** (a) By direct combination. 1. Light one end of about 3 in. of magnesium ribbon. Collect the product (white  $\text{MgO}$ ) in a clean evaporating dish. 2. Add a little water; notice the effect on litmus. 3. Upon a fragment of porcelain, held as in 41, 10, heat a little piece of cadmium steadily in the outer blow-pipe flame. Brown  $\text{CdO}$  is produced. Recall 8, 5, 17, 6, 27, 29, 5, 33, 5, 41, 10. (b) By heating certain salts: 4. Heat in a test tube 1 gm. of copper carbonate ( $\text{CuCO}_3$ ); dark brown  $\text{CuO}$  is left; what gas escapes? Prove it. 5. Heat in a glass tube closed at one end  $\frac{1}{2}$  gm. of mercuric nitrate ( $\text{Hg}(\text{NO}_3)_2$ ); red  $\text{HgO}$  is left; what is the result of continuing to heat the  $\text{HgO}$ ? See 6, 1. 6. Heat in a test tube 1 gm. of ferrous oxalate. Notice how the hot  $\text{FeO}$  when poured out becomes red hot and changes to a rust color. It oxidizes to  $\text{Fe}_2\text{O}_3$ . What gases escape? Suppose

you collect a test tube full over dilute NaOH, which gas would go into the test tube? Recall also **25** and **36**, 3. (c) From the hydroxid: 7. To solution of  $\text{CuSO}_4$  add a few drops of NaOH solution; pale blue copper hydroxid, changing to dark CuO on warming. 8. To solution of  $\text{Fe}_2\text{Cl}_6$  add  $\text{NH}_4\text{OH}$  to alkaline reaction; filter and heat some of the precipitate on a fragment of porcelain; red brown  $\text{Fe}_2\text{O}_3$  is left. Recall **7**, 8, **33**, 2, **37**, 2.

Equations: (1)  $\text{Mg} + \text{O} =$  . (2)  $\text{MgO} + \text{H}_2\text{O} =$  . (3)  $\text{Cd} + \text{O} =$  . (4)  $\text{CuCO}_3 =$  + . (5) Complete reaction uncertain; the brown fumes are oxids of Nitrogen. (6)  $\text{FeC}_2\text{O}_4 = \text{FeO} +$  + . (7)  $\text{CuSO}_4 + ? \text{NaOH} = \text{CuO}_2\text{H}_2 +$  .  $\text{CuO}_2\text{H}_2 = \text{CuO} +$  . (8)  $\text{Fe}_2\text{Cl}_6 + ? \text{NH}_4\text{OH} = \text{Fe}_2\text{O}_6\text{H}_6 + ? \text{NH}_4\text{Cl}$ .  $\text{Fe}_2\text{O}_6\text{H}_6 =$  + .

### *Hydroxids (Compounds with OH).*

**43.** Recall **9**, 7, **42**, 2, **23**, 3 and 5, **42**, 7 and 8, also **11**.

### *Sulfids (Compounds with S).*

**44.** Recall **5**, 2, **6**, 2, **18**, 1, 4, 6, **24**, 1. Put a little  $\text{Sb}_2\text{S}_3$  upon charcoal and heat it with the oxidizing flame; notice the odor of  $\text{SO}_2$  and the formation of some whitish  $\text{Sb}_2\text{O}_3$ . Recall **11**, 4, **17**, 8, **28**, 4, **34**, 3.

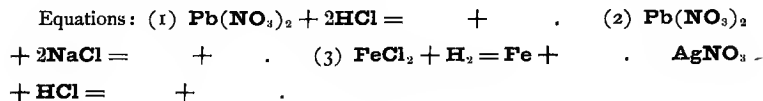
### *Chlorids (Compounds with Cl).*

**45.** Recall **7**, 2, 3, 9, **8**, 1, **13**, 1, **36**, 1, **39**, 1, 2. 1. To a solution of  $\text{Pb}(\text{NO}_3)_2$  or lead acetate add dilute HCl, drop by drop, till no more precipitate ( $\text{PbCl}_2$ ) is formed. 2. The same as 1, using a solution of any soluble chlorid ( $\text{NaCl}$  or  $\text{NH}_4\text{Cl}$ ) instead of HCl. This method applies to



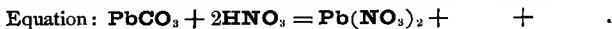


silver and mercurous salts also, and distinguishes them with lead from other metals. [3. Arrange for the preparation of H, using dilute  $\text{H}_2\text{SO}_4$  and Zn in either a s. n. test tube or an Erlenmeyer flask. But look over 7, and remember particularly about *knowing* the H is pure before applying a light. Put in a piece of glass tubing about  $\frac{1}{2}$  gm. of  $\text{FeCl}_2$  (ferrous chlorid), connect with the H generator, heat the tube under the  $\text{FeCl}_2$ ; hold moistened litmus paper at the mouth of the tube; this proves an acid. Dip a clean glass rod in  $\text{AgNO}_3$  solution and hold at the mouth of the tube; HCl, if present, will form white curdy  $\text{AgCl}$ . Compare with 7, 8.] Recall 14, 1, 4, 16, 4.



#### *Nitrates (Compounds with $\text{NO}_3$ ).*

46. Recall 26, 1, 28, 1. 1. Dissolve 2 gm.  $\text{PbCO}_3$  in as little hot dilute  $\text{HNO}_3$  as will dissolve it;  $\text{Pb}(\text{NO}_3)_2$  is formed which crystallizes from a strong solution on cooling. 2. Sprinkle a little  $\text{KNO}_3$  upon glowing charcoal; deflagration; it is an oxidizing agent. Recall 25, 42, 5.



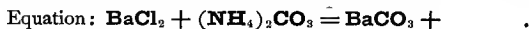
#### *Sulfates (Compounds with $\text{SO}_4$ ).*

47. Look up carefully the references under 20; see how many sulfates you find and how they were produced. To a solution of  $\text{Pb}(\text{NO}_3)_2$  add dilute  $\text{H}_2\text{SO}_4$  as long as it gives a precipitate —  $\text{PbSO}_4$ . By a similar method could be made  $\text{BaSO}_4$ ,  $\text{SrSO}_4$ , and less easily  $\text{CaSO}_4$ .



***Carbonates (Compounds with CO<sub>2</sub>).***

**48.** Recall **39, 42, 4**. To a solution of BaCl<sub>2</sub> add solution of (NH<sub>4</sub>)<sub>2</sub>CO<sub>3</sub>. White precipitate of BaCO<sub>3</sub>. This method applies to many of the metals, but in some cases the carbonate of the metal soon changes to the hydroxid; e. g. iron. See whether the BaCO<sub>3</sub> will dissolve in HCl—distinction from sulfates (**20, 4, 5**). What gas always escapes when a carbonate dissolves in an acid?

***The Formation of Insoluble Compounds—Precipitation.***

**49.** This occurs upon mixing soluble substances, when they, by the exchange of their corresponding parts, can form a compound which is insoluble. The corresponding parts are metals, or acid radicals. We have had many illustrations in the foregoing paragraphs; e. g. **6, 3, 8, 2, 17, 8, 18, 4, 6, 20, 4, 5, 24, 1, 32, 1, 2, 3, 35, 2, 36, 1, 42, 7, 8, 45, 1, 47, 48**. In the following experiments (1) note the soluble compounds given; (2) note the possible products, by double decomposition—exchange of corresponding parts; (3) consult Table of Solubility to ascertain whether either of the products is insoluble; (4) make prediction; (5) finally mix the substances and see whether their action corresponds to your prediction. Thus (1) lead nitrate and potassium iodid (2) might give potassium nitrate and lead iodid; (3) potassium nitrate is soluble, lead iodid is (nearly) insoluble; (4) there will be then a precipitate of lead iodid, while potassium nitrate remains in solution. (5) Try them. Will there be a precipitate upon mixing solutions of 1. Pb(NO<sub>3</sub>)<sub>2</sub> and KI? 2. Na<sub>2</sub>SO<sub>4</sub> and KI? 3. BaCl<sub>2</sub> and Na<sub>2</sub>SO<sub>4</sub>? 4. CuSO<sub>4</sub> and Na<sub>2</sub>SO<sub>4</sub>? 5. CuSO<sub>4</sub> and NaCl? 6. Na<sub>2</sub>CO<sub>3</sub> and Pb(NO<sub>3</sub>)<sub>2</sub>?







**Table of Solubility for 49.**

	Carbonate.	Chlorid.	Iodid.	Nitrate.	Sulfate.
Barium	I	S	S	S	I
Copper (ic)	I	S	S	S	S
Lead	I	I*	I*	S	I
Potassium	S	S	S	S	S
Sodium	S	S	S	S	S

Solubility and insolubility are of great importance in chemical operations because:—(1) Substances are often separated from one another by the application of solvents, e. g. salt from silicic oxid by water, in **24**, 3. (2) Substances are often made, in the laboratory and on the manufacturing scale, by precipitation; see Part IV, **8**. (3) Qualitative analysis depends largely upon a knowledge of the unlike behavior of different substances with precipitating and dissolving reagents; e. g. how could you distinguish a solution of either a silver, lead or mercurous salt from other metals? and also how could you distinguish a salt of lead from either of the other two? See last part of this section. For another illustration see **18**, 6. (4) Quantitative analysis often proceeds by converting substances into insoluble compounds and weighing; e. g. all of the chlorin in a solution of HCl may be combined by the addition of  $\text{AgNO}_3$ , into insoluble  $\text{AgCl}$ , from the weight of which the Cl and the HCl may readily be calculated.

It may be convenient for the learner to remember that, speaking generally, the following compounds are *soluble in water*: compounds of sodium, potassium, ammonium; nitrates; acetates; chlorids (except of  $\text{Ag}$ ,  $\text{H}_2\text{Pb}$ ); sulfates (except of  $\text{Pb}$ ,  $\text{Ba}$ ,  $\text{Sr}$ ,  $\text{Ca}$ ); bi or acid carbonates. Generally

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\* Soluble in hot water, or large quantities of cold water.

*insoluble* are metals and oxids of metals (except in a few cases with chemical change); hydroxids; sulfids; normal carbonates; phosphates; (except compounds of Na, K,  $\text{NH}_4$  and except the hydroxids and sulfids of Mg, Ca, Sr, Ba). Each compound, however, has its own peculiar solubility, which can only be known by experience and consultation of the more complete tables of solubility.

***Action of Metals upon Solutions of Metallic Salts.***

[50. In the following experiments two facts are to be verified in each case—first, that the metal in solution is partly or wholly deposited, evidenced by sight of the deposit; second, that the metal at the outset uncombined has partly entered into the solution, shown by testing the solution with some characteristic reagent before the change, and then testing it, or another portion of it, after the change.

1. Into each of two test tubes put about 8c.c. of  $\text{AgNO}_3$  solution; to one tube add a globule of Hg;  $\text{HgNO}_3$  and Ag are produced. After five to ten minutes look for the Ag, then add KI to both tubes. AgI is yellow, HgI is greenish. If any HgI is produced, it shows that Hg had gone into the liquid as a soluble compound. 2. Similarly  $\text{HgCl}_2 + \text{Cu} = \text{CuCl}_2 + \text{Hg}$ . Test with  $\text{NH}_4\text{OH}$ ; a blue liquid shows copper. 3.  $\text{Cu}(\text{NO}_3)_2 + \text{Pb} = \text{Pb}(\text{NO}_3)_2 + \text{Cu}$ . Test with dilute  $\text{H}_2\text{SO}_4$ ; Pb will give white  $\text{PbSO}_4$ . 4.  $\text{Pb}(\text{NO}_3)_2 + \text{Zn} = \text{Zn}(\text{NO}_3)_2 + \text{Pb}$ . Test with  $\text{K}_3\text{FeC}_6\text{N}_6$  (ferricyanid of potassium); ferricyanid of Zn is brownish.

Notice that Zn displaces Pb; Pb displaces Cu; Cu displaces Hg; Hg displaces Ag.]





*To the Pupil.*

At the conclusion of this part of your work you will be expected to identify the substances used ; at least those included in the following list. You will have no difficulty in doing this, if you follow this plan :—

1. Define.
2. Classify.
3. Distinguish.

1. Define, by writing against each substance a very brief description — viz. “a blue liquid,” “a yellow, brittle solid.”

2. Classify, by arranging in groups according to obvious properties ; this can be done in many ways — solids, liquids ; colored, not colored ; with odor, without odor ; affecting litmus, not affecting litmus ; metals, not metals, etc.

3. Distinguish. If you classify well, the groups distinguish themselves, and you have only to find some physical or chemical difference between a few of the same group. Thus you can easily tell copper from iron or zinc or cadmium or tin by its appearance ; you can tell cadmium from tin, not by its appearance but by its brown oxid when heated with the blow-pipe or by the fact that it dissolves in  $\text{HNO}_3$ , while tin does not.

## LIST.

## PROPERTIES.

copper sulfate, crystals  
copper sulfate, solution  
ferric oxid  
sulfur, broken roll  
sulfur, flowers  
copper, foil  
copper, sheet  
iron wire

LIST.	PROPERTIES.
mercuric oxid	
zinc dust	
zinc, granulated	
hydrochloric acid, conc.	
hydrochloric acid, dil.	
sulfuric acid, conc.	
sulfuric acid, dil.	
potassic chlorate	
sugar	
starch	
lime	
lead oxid	
carbonate of soda, cryst.	
calcic chlorid, fused	
"chlorid of lime "	
hydrogen peroxid	
permanganate of potassium, sol.	
sulfate of lime	
common salt, fine	
rock salt	
bromid of potassium	
iron sulfid	
iodin	
iodid of potassium	
paraffin	
water	
tin	
cadmium	
baric chlorid, sol.	
lead carbonate	
silver nitrate, sol.	
manganese dioxid	
ferric chlorid, sol.	
alcohol	
acetic acid	
sulfate of soda, sol.	
lead nitrate, sol.	
sulfite of soda	
ammonium chlorid	







LIST.	PROPERTIES.
lime water	
ammonia, sol.	
ammonium nitrate	
nitric acid, conc.	
nitric acid, dil.	
potassium nitrate	
bisulfid of carbon	
antimony	
sulfid of antimony	
silicate of soda, sol.	
charcoal	
bone black	
carbonate of lime	
caoutchouc	
lycopodium	
magnesium	
carbonate of copper	
iron oxalate	
How could you distinguish copper from charcoal?	
How could you distinguish solution of $\text{NH}_4\text{Cl}$ from solution of $\text{NaCl}$ ?	
How could you distinguish solution of $\text{NaCl}$ from $\text{Na}_2\text{SO}_4$ ?	
How could you distinguish fine salt from ferrous sulfid?	
How could you distinguish fine salt from lead carbonate?	
How could you distinguish dilute $\text{HCl}$ from $\text{H}_2\text{SO}_4$ ?	
How could you distinguish bisulfid of carbon from anything else?	
How could you distinguish O from H?	
How could you distinguish N from H?	
How could you distinguish $\text{HCl}$ gas from $\text{NH}_3$ gas?	
How could you distinguish $\text{SO}_2$ from $\text{H}_2\text{S}$ ?	
How could you distinguish $\text{CaCO}_3$ from $\text{Na}_2\text{CO}_3$ ?	

## PART II.

### CHEMICAL COMPUTATIONS

Atomic weights.

CLASS I. Constants from experimental data

1. Vapor Density.
2. Molecular Weight of Volatile Substances.
3. Molecular Weight of Solids.
4. Atomic Weight.
5. (a) Formula of Volatile Substances.  
(b) Formula of Solids.
6. Valence.
7. Equations.

CLASS II. Practical calculations from accepted constants.

8. Molecular or Formula Weight.
9. Vapor Density.
10. Weight of gas, volume given.
11. Volume of gas, weight given.
12. Correction of gaseous volumes.
  - (a) For Temperature.
  - (b) For Pressure.
  - (c) For Tension.
13. Percentage, from Formula.
14. Weight of constituent of compound.
15. Computations from Equation.
  - (a) Weight.
  - (b) Volume.
16. In the Metric System.
  - (a) Weight of liquids from measure.
  - (b) Measure of liquids from weight.
17. Normal Solutions.

Examples applying these principles.





A Table of Atomic Weights (O=16). Revised to January 1, 1895, by F. W. Clarke.

NAME.	SYMBOL.	ATOMIC WEIGHT.	NAME.	SYMBOL.	ATOMIC WEIGHT.
Aluminum .....	<b>Al</b> .....	27.	Neodymium .....	<b>Nd</b> .....	140.5
Antimony .....	<b>Sb</b> .....	120.	Nickel .....	<b>Ni</b> .....	58.7
Arsenic .....	<b>As</b> .....	75.	Nitrogen .....	<b>N</b> .....	14.03
Barium .....	<b>Ba</b> .....	137.43	Osmium .....	<b>Os</b> .....	190.8
Bismuth .....	<b>Bi</b> .....	208.	Oxygen .....	<b>O</b> .....	16.
Boron .....	<b>B</b> .....	11.	Palladium .....	<b>Pd</b> .....	106.5
Bromin .....	<b>Br</b> .....	79.95	Phosphorus .....	<b>P</b> .....	31.
Cadmium .....	<b>Cd</b> .....	112.	Platinum .....	<b>Pt</b> .....	195.
Caesium .....	<b>Cs</b> .....	132.9	Potassium .....	<b>K</b> .....	39.11
Calcium .....	<b>Ca</b> .....	40.	Praseodymium ..	<b>Pr</b> .....	143.5
Carbon .....	<b>C</b> .....	12.	Rhodium .....	<b>Rh</b> .....	103.
Cerium .....	<b>Ce</b> .....	140.2	Rubidium .....	<b>Rb</b> .....	85.5
Chlorin .....	<b>Cl</b> .....	35.45	Ruthenium .....	<b>Ru</b> .....	101.6
Chromium .....	<b>Cr</b> .....	52.1	Samarium .....	<b>Sm</b> .....	150.
Cobalt .....	<b>Co</b> .....	59.5	Scandium .....	<b>Sc</b> .....	44.
Columbium .....	<b>Cb</b> .....	94.	Selenium .....	<b>Se</b> .....	79.
Copper .....	<b>Cu</b> .....	63.6	Silicon .....	<b>Si</b> .....	28.4
Erbium .....	<b>Er</b> .....	166.3	Silver .....	<b>Ag</b> .....	107.92
Fluorin .....	<b>F</b> .....	19.	Sodium .....	<b>Na</b> .....	23.05
Gadolinium .....	<b>Gd</b> .....	156.1	Strontium .....	<b>Sr</b> .....	87.66
Gallium .....	<b>Ga</b> .....	69.	Sulfur .....	<b>S</b> .....	32.06
Germanium .....	<b>Ge</b> .....	72.3	Tantalum .....	<b>Ta</b> .....	182.6
Glucinum .....	<b>Gl</b> .....	9.	Tellurium .....	<b>Te</b> .....	125.
Gold .....	<b>Au</b> .....	197.3	Terbium .....	<b>Tb</b> .....	160.
Hydrogen .....	<b>H</b> .....	1.008	Thallium .....	<b>Tl</b> .....	204.18
Indium .....	<b>In</b> .....	113.7	Thorium .....	<b>Th</b> .....	232.6
Iodin .....	<b>I</b> .....	126.85	Thulium .....	<b>Tu</b> .....	170.7
Iridium .....	<b>Ir</b> .....	193.1	Tin .....	<b>Sn</b> .....	119.
Iron .....	<b>Fe</b> .....	56.	Titanium .....	<b>Ti</b> .....	48.
Lanthanum .....	<b>La</b> .....	138.2	Tungsten .....	<b>W</b> .....	184.9
Lead .....	<b>Pb</b> .....	206.95	Uranium .....	<b>U</b> .....	239.6
Lithium .....	<b>Li</b> .....	7.02	Vanadium .....	<b>V</b> .....	51.4
Magnesium .....	<b>Mg</b> .....	24.3	Ytterbium .....	<b>Yb</b> .....	173.
Manganese .....	<b>Mn</b> .....	55.	Yttrium .....	<b>Yt</b> .....	89.1
Mercury .....	<b>Hg</b> .....	200.	Zinc .....	<b>Zn</b> .....	65.3
Molybdenum .....	<b>Mo</b> .....	96.	Zirconium .....	<b>Zr</b> .....	90.6

## CHEMICAL COMPUTATIONS.

CLASS I. — *Based upon experimental data, and illustrating methods of establishing Atomic Weights, Formulas, and Equations.*

1. The vapor density of a gas is its weight in comparison with the weight of its own volume of hydrogen.

*Note.*—Chemists have generally accepted the view stated in the "Theory of Avogadro"; viz. that equal volumes of gases contain the same number of molecules. According to this view the number obtained for Density as above would also be the weight of 1 molecule (of the gas under discussion) in comparison with 1 molecule of hydrogen.

METHOD: Weighing a definite volume of the gas.

ILLUSTRATION: The weight of 1 L of ammonia gas is .7616 gm.; the weight of 1 L of hydrogen is .0896 gm.  $.7616 \div .0896 = 8.5$ . The vapor density of ammonia is 8.5, i. e. it is 8.5 times as heavy as the same volume of hydrogen under standard conditions.

2. The molecular weight of a gas is its weight compared with the weight of half its own volume of hydrogen—i. e. twice its density.

*Note.*—According to Avogadro's Theory, molecular weight, as obtained above, would be the weight of 1 molecule (of the gas under discussion) in comparison with the half-molecule of hydrogen.

DATUM: Vapor density.

METHOD: Multiplication by 2. When the density is for any reason stated in terms of air, the multiplication is by  $14.44 \times 2$ .

ILLUSTRATION: The vapor density of O, N, CO<sub>2</sub> is 16, 14, 22, respectively; the molecular weight is 32, 28, 44.

3. The molecular weight of a non-volatile solid is in some cases determined by certain physical methods, for which reference must be made to a larger treatise.

E. G. — See "*Solutions*," by W. Ostwald, upon Vapor Pressures and Freezing Points.







4. The atomic weight of an element is the relative weight (\*H being 1 or O being 16) in which the element enters into chemical changes and forms part of chemical compounds. It is determined by as many of the following methods as can be applied:—

**A. DATA:** Percentage composition and vapor density of a large number of gaseous compounds in which the element occurs.

**METHOD:** The vapor densities give the molecular weights of the compounds; the percentages give the quantity of the element, whose atomic weight is sought, in those molecular weights. The greatest common divisor of the numbers expressing the quantity of the element is the atomic weight.

**ILLUSTRATION:** Oxygen occurs in water, carbon dioxide, sulfur trioxide, etc. Their vapor densities are 9, 22, 40, and their molecular weights, 18, 44, 80, respectively. Oxygen is 88.89% of the water, or  $\frac{16}{18}$ ; 72.72% of carbon dioxide, or  $\frac{32}{44}$ ; and 60% of sulfur trioxide, or  $\frac{48}{80}$ . 16 is considered the atomic weight of oxygen.

**B. DATA:** The weight of the element which displaces 1 part of hydrogen from a compound.

**C. DATA:** The weight of the element which combines with 35.45 parts of chlorine.

**METHOD:** The number obtained by either **B** or **C** is the same, and is either the atomic weight, or a sub-multiple of the atomic weight. This method is accurate for the ratio.

**ILLUSTRATION:** 107.9 parts of silver displace 1 of hydrogen. 103.47 of lead combine with 35.45 of chlorine. The atomic weights of silver and lead are either 107.9 and 103.47, or two or three, etc., times these numbers. See **D**.

---

\* *Note.*—Since the ratio of O to H is not at present considered exactly 16 to 1, but 15.88 to 1, two slightly differing systems are in use—one in which the computations are to O as 16 (H = 1.008), and the other in which H is taken as 1 (O = 15.88). In the table on page 61 O = 16 is the standard.

**D. DATA:** The specific heat of the element, and the number 6.4\*, which is called the atomic heat, and which is approximately the same for all solid elements.

**METHOD:** 6.4 divided by the specific heat of the element gives its atomic weight. This result is approximate only, but settles the question left undecided by **B** and **C**.

**ILLUSTRATION:** The specific heat of silver is .0559.  $6.4 \div .0559 = 114$ ; and the atomic weight of silver is 107.9, and not a multiple of 107.9. The specific heat of lead is .0315.  $6.4 \div .0315 = 203$ ; and the atomic weight of lead is not 103.47, but twice as much, or 206.95.

**5. A.** A chemical formula of a gas or volatile substance expresses the proportional composition of the substance and a total weight equaling twice its vapor density.

**DATA:** The vapor density of the compound, the proportions of the constituent elements in the compound, and the atomic weights of the elements.

**METHOD:** Obtain molecular weight: compute the parts of it indicated by the proportion of each element in the compound: these numbers divided by the atomic weights of the corresponding elements give the number of atoms of each element.

**ILLUSTRATION:** Water has a vapor density of 9 and molecular weight of 18, and consists of 11.14% hydrogen and 88.86% oxygen. The atomic weights of H and O are 1 and 16 respectively.  $11.14\%$  of 18 =  $2 \div 1 = 2$  or  $H_2$ .  $88.86\%$  of 18 =  $16 \div 16 = 1$  or O. Formula,  $H_2O$ .

**B.** A chemical formula of a non-volatile substance usually expresses the proportional composition of the substance in the simplest manner consistent with the weight meaning implied in the symbols.

**DATA:** Proportional composition of the substance, and the atomic weights of the elements.

**METHOD:** Divide the percentage numbers by the atomic weights respectively: results represent relative number of atoms: these results are usually fractional, and are to be reduced to integers.

---

\* Variously estimated; — from 6.1 to 6.5.





## ILLUSTRATION :

Analysis gives, C  $35.82 \div 12 = 2.985, \div 2.985 = 1, \times 4 = 4.$

H  $4.479 \div 1 = 4.479, \div 2.985 = 1\frac{1}{2}, \times 4 = 6.$

O  $59.7 \div 16 = 3.731, \div 2.985 = 1\frac{1}{2}, \times 4 = 5.$

99.99+

Formula is  $C_4H_6O_5$ .

(This method is now being modified by the new methods for determining molecular weight of solids suggested in 3.)

**6.** The Valence (Equivalence, Quantivalence) of an element is practically the number of times 1 of H, 35.4 of Cl, or 8 of O, with which its own atomic weight unites.

The idea of valence depends upon the correct determination of formulas and atomic weights. Thus from the formula  $HCl - 1$  to 35.4 — we say the valence of Cl is 1. Its atomic weight combines with the atomic weight of the standard — H. From the formula  $NaCl - 23$  to 35.4 — we say the valence of Na is 1; from the formula  $BaCl_2 - 137$  to 70.8 — we say the valence of Ba is 2.  $SnO$  and  $SnO_2$  seem to show the valence of Sn as sometimes 2, sometimes 4. Valence is indicated —  $H', Cl', Ba'', Sn^{II, IV}$

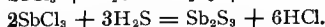
Valence is useful in assisting the memory of formulas, or in conjecturally writing unknown formulas; but it must be remembered that, experimentally, valence is an inference from formulas, not formulas from valence.

**7.** A chemical equation indicates, in their relative proportions, all the materials which enter into and proceed from a chemical change.

**DATA:** The only complete data for an equation are the weights of all the factors and products. When, however, all the factors and products are known, and their formulas, a correct equation can usually be written.

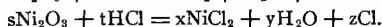
**A. METHOD:** Inspection. Write the smallest number of whole molecules which balance.

ILLUSTRATION:  $\text{SbCl}_3 + \text{H}_2\text{S} = \text{Sb}_2\text{S}_3 + \text{HCl}$ .



B. METHOD: Assume coefficients, and form equations between the coefficients of the same element. Solve.

ILLUSTRATION:  $\text{Ni}_2\text{O}_3 + \text{HCl} = \text{NiCl}_2 + \text{H}_2\text{O} + \text{Cl}$ .



$2s = x$ ,  $3s = y$ ,  $t = 2y$ ,  $t = 2x + z$ . Assume  $s = 1$ , then  $x = 2$ ,  $y = 3$ ,  $t = 6$ ,  $z = 2$ , and the equation is  $\text{Ni}_2\text{O}_3 + 6\text{HCl} = 2\text{NiCl}_2 + 3\text{H}_2\text{O} + 2\text{Cl}$ , or  $\text{Cl}_2$ .

CLASS II. — *Practical calculations in which Atomic Weights, Formulas, and Equations are accepted as established facts.*

8. To compute molecular weight or formula weight.

DATA: The formula of the compound; atomic weights, by implication.

METHOD: The sum of the atomic weights in the formula is the molecular or formula weight.

ILLUSTRATION: The molecular weight of  $\text{H}_2\text{SO}_4 = 2 + 32 + 64 = 98$ .

9. To compute vapor density (of a gas or volatile substance).

DATA: Formula.

METHOD: Take half the molecular weight.

ILLUSTRATION: The density of alcohol vapor ( $\text{C}_2\text{H}_5\text{OH}$ ) is

$$\frac{24 + 5 + 16 + 1}{2} = 23.$$

This gives hydrogen density; if air density is desired, the result is divided by 14.44.

10. To compute the weight of a given volume of gas.

DATA: Formula or density of gas, and number of litres; also the weight of 1 L. of hydrogen (.0896 grammes).

METHOD: Multiply the density of the gas by .0896 gr., and the product by the volume of the gas in litres.  $W = V. D. \times .0896 \times L.$

ILLUSTRATION: The weight of 25 L. of oxygen, V. D. = 16, is  $16 \times .0896 \times 25 = 35.84$  gr.







**11.** To compute the volume of a given weight of gas.

DATA: Formula or density of gas and total weight; weight of 1 L. of H.

METHOD: Divide the given weight by the density times .0896 gr.

$$V. \text{ or } L. = \frac{W}{V. D. \times .0896}.$$

ILLUSTRATION: How many litres are 17 grammes of  $\text{CO}_2$ ?

$$\frac{17}{22 \times .0896} = 8.62 \text{ L.}$$

**12.** To correct measured volume of gases for temperature, pressure, and tension of water vapor.

The volume of all gases is referred to  $0^\circ$ , 760m.m. and a condition of dryness. It is not practically convenient to measure under exactly the above conditions; so gases are measured at convenient temperature and pressure and corrections are made.

**A. FOR TEMPERATURE:** Multiply the measured volume by 273 and divide by  $273 +$  (or  $-$ ) the number of degrees above (or below)  $0^\circ\text{C}$ .

$$V = \frac{V^1 \times 273}{273 \pm t}.$$

**B. FOR PRESSURE:** Multiply the measured volume by the number of millimetres of barometric pressure, and divide by 760.

$$V = \frac{V^1 \times P}{760}.$$

Either correction may be made upon the result of the other, or both may be combined in one operation.

**C.** The tension of water vapor is usually applied from a table and subtracted from  $P$  the observed pressure (at  $15^\circ\text{C}$  the correction is 12.7m.m.).

ILLUSTRATION: A gas measures over water, 450c.c. at 750m.m. and  $15^\circ\text{C}$ . What is its normal volume?

$$V = 450 \times \frac{273}{288} \times \frac{737.3}{760} = 413.8.$$

**13.** To compute percentage composition of any compound from its formula.

DATA: Formula, implying atomic weights.

METHOD: Take such part of 100 as the weight of any constituent, expressed by the formula, is of the whole molecular weight.

ILLUSTRATION: The percentage composition expressed by  $\text{Na}_2\text{CO}_3$  is computed as follows:—

$$\begin{array}{rcl}
 \text{Na}_2 & = 46 & \text{Na}_2 = \frac{46}{106} \text{ of } 100 = 43.39\% \\
 \text{C} & = 12 & \text{C} = \frac{12}{106} \text{ of } 100 = 11.32\% \\
 \text{O}_3 & = 48 & \text{O}_3 = \frac{48}{106} \text{ of } 100 = 45.28\% \\
 \hline
 & 106 &
 \end{array}$$

**14.** To compute the weight of any constituent in a given weight of a compound.

DATA: The formula of the compound and the given weight.

METHOD: Take such part of the given weight as the weight of any constituent, expressed by the formula, is of the whole molecular weight.

ILLUSTRATION: How many grammes of oxygen in 200 grammes of  $\text{KClO}_3$ ?

$$\begin{array}{rcl}
 \text{K} & = 39. & \\
 \text{Cl} & = 35.5. & \\
 \text{O}_3 & = 48. & \text{O}_3 = \frac{48}{122.5} \text{ of } 200 \text{ or } 78.36 \text{ grammes.} \\
 \hline
 & 122.5 &
 \end{array}$$

**15.** To compute from equations.

**A. Weight.**

DATA: The equation and the weight of one of the compounds.

METHOD: The weight of any compound is to the weight of any other compound as the relative weight in the equation of the first is to the relative weight of the other.  $W:w = M:m$ .

ILLUSTRATION:  $\text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$ .

What weight of  $\text{H}_2\text{SO}_4$  is needed with 15 gr. of  $\text{Na}_2\text{CO}_3$ ?  $106:98 = 15:x$   
 $= 13.8$ .

How much  $\text{Na}_2\text{SO}_4$  can be made from 15 gr. of  $\text{Na}_2\text{CO}_3$ ?  $106:142 = 15:x = 20$ .

106, 98, 142 are the relative weights expressed by  $\text{Na}_2\text{CO}_3$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{Na}_2\text{SO}_4$ .

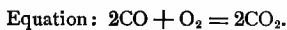




**B. Volume.**

When the equation involves gases each formula represents the same volume. (This is an inference from **1**, **2** and **5**). Thus  $\text{CH}_4 + 2\text{O}_2 = \text{CO}_2 + 2\text{H}_2\text{O}$ , and volume computations can be made; but  $\text{O}_2$ ,  $\text{H}_2$ ,  $\text{N}_2$ ,  $\text{Cl}_2$  are the formulas (not O, H, N, Cl) of simple gases.

ILLUSTRATION: How many litres of O are required to burn 1 litre of CO?



4 vols.    2 vols.    4 vols.

$4 : 2 :: 1 : x$ ,  $x = \frac{1}{2}$ .    *Ans.*

**16.** In the metric system of weights and measures the relations between weight, measure, and specific gravity are very convenient. 1 cubic centimeter of water (by measure) = 1 gram (by weight); and the specific gravity of water is 1.

To calculate weight from measure, of liquids.

DATA: Measure and specific gravity.

METHOD: Multiply the number of cubic centimeters by specific gravity: the result is grams.

ILLUSTRATION: 400c.c. of alcohol, specific gravity .80, weighs  $400 \times .80 = 320$  grams.

**B. To calculate measure from weight, of liquids.**

DATA: The weight and specific gravity.

METHOD: Divide the number of grams by the specific gravity.

ILLUSTRATION: 370 gm. of  $\text{H}_2\text{SO}_4$ , specific gravity 1.85 measures  $370 \div 1.85 = 200\text{c.c.}$

**17. Normal solutions.**

A normal solution of a chemical reagent means primarily the formula weight of the substance, in grams, with enough water to measure 1 litre (1,000c.c.). Thus normal sodium hydroxid (NaOH) would contain 40 gm. in 1 litre of

solution, or 40 mgm. in 1 cubic centimeter. In the case, however, of diacid bases, like  $\text{Ca}(\text{OH})_2$ , and dibasic acids, like  $\text{H}_2\text{SO}_4$ , one-half the formula weights, 37 and 49, would be used, thus giving 1c.c. the same neutralizing power as 1c.c. of  $\text{NaOH}$  or  $\text{HCl}$ ; and other reagents are made up on the same principle.

In practice much more dilute solutions are used —  $\frac{1}{10}$  and  $\frac{1}{20}$  normal — which are often written  $\frac{N}{10}$  and  $\frac{N}{20}$ .

The strength of exact solutions is often indicated fractionally, the numerator denoting the weight of reagent in grams, and the denominator the total measure of the solution in centimeters; thus  $\frac{5.6}{1000}$  might represent normal  $\text{KOH}$ .

The principles embodied in the problems given above find many applications. A few are given below. The teacher can multiply such problems indefinitely and give them out upon numbered cards.

The pupil is advised to indicate all the arithmetical work in each operation before performing it; also, he should, as soon as able, perform multiplications and divisions by means of logarithms. After the principle is understood, the work is much simplified by the use of Tables, e. g. Adriance's Laboratory Calculations.

1. Calculate the molecular weights of  $\text{HCl}$ ,  $\text{H}_2\text{S}$ ,  $\text{CO}$ ,  $\text{CO}_2$ ,  $\text{NH}_3$ ,  $\text{N}_2\text{O}$ ,  $\text{SO}_2$ .

2. Calculate the vapor densities of each of the above gases.

3. Which of them are heavier than air?

4. Calculate the percentage composition of  $\text{KBr}$ ,  $\text{NaBr}$ ,  $\text{NH}_4\text{Br}$ ,  $\text{NiBr}_2$ ,  $\text{MgBr}_2$ ,  $\text{LiBr}$ , and arrange them in the order of most or least  $\text{Br}$ .







5. Calculate the true volume of 300c.c. of oxygen, measured at  $20^{\circ}\text{C}$  and 740m.m. pressure.

6. What is the normal volume of 300c.c. of oxygen measured at  $-20^{\circ}\text{C}$  and 780m.m.?

7. Compute the weight of 1 litre of each of the gases mentioned in 1.

8. If 9 parts of Al combine with 35.45 of Cl and the specific heat of Al is .2253, what is the atomic weight of Al?

9. What is the per cent. of iron in  $\text{Fe}_4(\text{P}_2\text{O}_7)_3, 9\text{H}_2\text{O}$ ?

10. What weight and volume of H is obtainable by 4 gm. of Zn?

11. Compute the formula of a gas found by experiment to consist of N, 63.63%, O, 36.36%, and to have vapor density 22.

12. How many centimeters of HCl solution, specific gravity 1.21, containing 42.4% HCl by weight, are needed for 4 litres of chlorine?  $4\text{HCl} + \text{MnO}_2 = \text{Cl}_2$ , etc.

13. Assign formula to a solid found by experiment to consist of Ca, 29.41%; S, 23.52%; O, 47.05%.

14. Na = 39.31%; Cl = 60.68%; assign formula.

15. See II, 15. (1) How much  $\text{H}_2\text{SO}_4$  with 10 gm.  $\text{Na}_2\text{CO}_3$ ? (2) How much  $\text{H}_2\text{SO}_4$  for 10 gm.  $\text{Na}_2\text{SO}_4$ ? (3) How much  $\text{H}_2\text{SO}_4$  for 10 gm.  $\text{CO}_2$ ? (4) How much  $\text{Na}_2\text{CO}_3$  with 10 gm.  $\text{H}_2\text{SO}_4$ ? (5) For 10 gm.  $\text{Na}_2\text{SO}_4$ ? (6) For 10 gm.  $\text{CO}_2$ ? (7) How much  $\text{CO}_2$  from 10 gm.  $\text{Na}_2\text{CO}_3$ ? (8) How much  $\text{CO}_2$  when 10 gm.  $\text{H}_2\text{SO}_4$  are used? (9) When 10 gm.  $\text{Na}_2\text{SO}_4$  are produced? (10) How much  $\text{Na}_2\text{SO}_4$  from 10 gm.  $\text{H}_2\text{SO}_4$ ? (11) From 10 gm.  $\text{Na}_2\text{CO}_3$ ? (12) When 10 gm.  $\text{CO}_2$  are produced?

16. In 2,000 lbs. of  $\text{NaNO}_3$ , 85% pure, how much Na?

17. How much  $\text{H}_2\text{SO}_4$  is needed to make 25 litres of  $\text{CO}_2$  from  $\text{CaCO}_3$ ?

18. Suppose a manufacturer of  $\text{H}_2\text{SO}_4$  to start with 100 lbs. of sulfur and to obtain the full theoretical yield in each case:

(1) What weight of O is needed to burn the S to  $\text{SO}_2$ ?

(2) What additional weight of O to convert the  $\text{SO}_2$  to  $\text{SO}_3$ ?

(3) What weight of  $\text{H}_2\text{O}$  to convert the  $\text{SO}_3$  to  $\text{H}_2\text{SO}_4$ ?

(4) And how much  $\text{H}_2\text{SO}_4$  is obtained?

19. What weight of  $\text{KClO}_3$  will yield the oxygen to fill two 300c.c. jars?

---

20. A student is experimentally determining the amount of Ba in 1 gm. of barium chlorid ( $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ ). He does this by converting all the Ba into  $\text{BaSO}_4$ , and weighing that. He obtains .96 gm.  $\text{BaSO}_4$ . How great is his error; i. e., the difference between the theoretical % of Ba in the barium chlorid and the % computed from his result?

21. 11c.c. of a  $\frac{1}{10}$  normal solution of nitrate of silver ( $\text{AgNO}_3$ ) is just sufficient for all the salt ( $\text{NaCl}$ ) in 250c.c. of a water which is being analyzed. How much  $\text{NaCl}$  is there in a litre of the water, and how much Cl?  $\text{AgNO}_3 + \text{NaCl} = \text{AgCl} + \text{NaNO}_3$ .

22. A substance believed to be cupric hydrid gives from analysis, copper 96.72%, hydrogen 3.28%; compare these results with the theoretical results from the formula  $\text{CuH}_2$ .

23. A solution of  $\text{H}_2\text{SO}_4$  intended to be  $\frac{N}{10}$  is found to contain .00524 gm. per cubic centimeter of acid; how much water must be added to 1 litre of this solution in order to make the resulting solution exactly  $\frac{N}{10}$ ?





## PART III.

### EXPERIMENTS IN QUANTITATIVE RELATIONS.

#### The Supplies.

1. The hydrogen set free by a definite weight of metal.
  - (a) Zinc and hydrochloric acid.
  - (b) Magnesium and sulfuric acid.
  - (c) Tin and hydrochloric acid.
2. The percentage of oxygen in chlorate of potassium ; the volume of a given weight of oxygen.
3. The percentage of water of crystallization in
  - (a) Baric chlorid.
  - (b) Copper sulfate.
  - (c) Sodium carbonate.
4. The weight of salt in solution at a given temperature.
5. The solubility of a salt at different temperatures.
6. The proportion by volume of oxygen in air.
7. The behavior of indicators.
8. The neutralizing power of an alkali ; a certain weight, and twice that weight.
  - (a) With hydrochloric acid.
  - (b) With sulfuric acid.
9. The relative neutralizing power of the hydroxids of potassium and sodium.
10.
  - (a) The quantity of sulfate, chlorid, and nitrate of sodium from a given weight of the carbonate.
  - (b) The calculated quantity of acids used.
  - (c) The basicity of the acids.

11. (a) The weight of copper precipitated by cadmium.  
 (b) The weight of lead precipitated by cadmium.  
 (c and d) The method applied to silver and mercury.
12. One-half the volume of hydrochloric acid gas is hydrogen.
13. The relative volumes of carbon monoxid and dioxid from the decomposition of ferrous oxalate.
14. Successive transformations of a lead compound with the weights of the products.  
 (a) The buff oxid from white lead.  
 (b) The yellow oxid from the buff oxid.  
 (c) The nitrate from the yellow oxid.  
 (d) The sulfate from the nitrate.

### Summary.

The experiments in Part III have been selected or devised with reference to reasonable accuracy without elaborate appliances. The pupil's own skill and good sense are important factors. He will commonly improve his work by trying again. Primary emphasis is laid upon the quantitative relations; but the pupil must comprehend every chemical change involved, and the reason for every process.

Each student will require the following articles; but it is not necessary to get a complete outfit for each of several pursuing the same course.

- 1 Stoddard's balance, or other balance weighing to 10 m.gr.
- 1 500c.c. flat-bottomed flask.
- 1 2-hole rubber stopper for flask.
- 1 bulbed ignition tube, 4"  $\times$   $\frac{1}{2}$ ".
- 1 1-hole rubber, or best velvet cork stopper for ignition tube.
- 2 royal berlin porcelain crucibles, 20c.c., or
- 2 royal berlin porcelain evaporating dishes, 30 to 50c.c.
- 1 dessicator (a covered dish with a platform conveniently of perforated tin; the dish containing about  $\frac{1}{2}$  in. of concentrated  $H_2SO_4$ ).
- 1 2-inch watch glass.







- 1 water bath (usually of copper; can be made by fitting a tin cover with suitable holes to an ordinary tin pan).
- 1 50c.c. burette, graduated to  $\frac{1}{10}$ .
- 1 stand for burette.
- 1 pipette to deliver 10c.c.
- 1 500c.c. side-necked Erlenmeyer filtering flask of stout glass.
- 1 solid rubber stopper for flask.
- 1 spherical separating funnel  $\frac{1}{2}$  or  $\frac{3}{8}$ " tube; soft glass tubing,  $\frac{1}{2}$  or  $\frac{3}{8}$ ".
- 1 or 2 stout pinch-cocks; rubber tubing, glass tubing, test tubes, etc., as in Part I.

The only chemicals required beyond the supply for Part I are indicators: litmus, cochineal, lacmoid, phenol-phthal-  
ein, etc., magnesium powder, pure, dry, white lead.

**1.** *The quantity of hydrogen set free from acids by definite weights of metals.*

\* APPARATUS. (See Fig. II.) It consists of a 500c.c. flask with a 2-hole rubber stopper, a side-necked test tube 6" x 1" with the neck bent down at a right angle and put through one hole in the stopper of the flask, and rubber stopper to fit; a gas jar holding about 400c.c.; a glass tube bent at right angles passes through the other hole of the stopper almost to the bottom of the flask and is connected to rubber tubing reaching to the bottom of the gas jar. At the beginning of the experiment the flask is filled with water to one inch of the stopper;

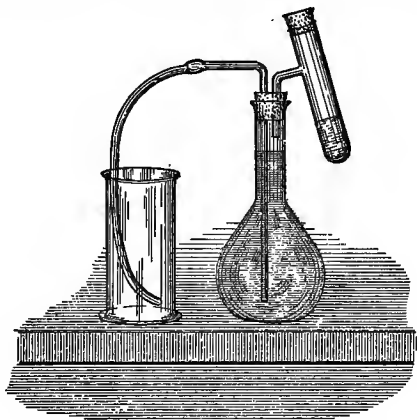


Fig. II.

\* Modified from Ramsey's "Chemical Theory."

acid of the specified strength is put in the s. n. test tube, making it about half full; all the stoppers and joints must be tight. The pupil should do two of either **A**, **B**, or **C**.

**A.** Zinc and HCl. Use dilute HCl (1 to 2 or 3), weigh out about 1 gm. of granulated zinc (not less than .9 gm. nor more than 1.1). Remove the stopper from the s. n. tube, drop in the zinc and instantly replace stopper. The hydrogen evolved will now force its volume of water from the flask into the gas jar. When all action is over, and the flask and all the apparatus have cooled to the temperature of the room, bring the liquids to the same level in flask and gas jar, detach rubber tube and let the water in it flow into the jar. Observe temperature of room, and barometric pressure, if possible. Now measure carefully the contents of the gas jar, making the best allowance you can for the glass tube in the flask.

#### DATA AND COMPUTATIONS.

Number of c.c. of H produced by	gm. of Zn	=
Corrected for temperature ( ° ), pressure ( . m.m.), and tension		=
Weight of H (computed; 1 litre = .0896 gm.)		=
Weight of Zn required for 1. of H, computed	}	=
Equivalent of Zn		
Accepted equivalent of Zn (= atomic weight ÷ 2)		=

**B.** Magnesium ribbon (.3 gm.) and  $\text{H}_2\text{SO}_4$  (1 to 10).

#### DATA AND COMPUTATIONS.

Number of c.c. of H produced by	gm. of Mg	=
Corrected for		=
Weight of H (computed)		=
Weight of Mg for 1 of H	}	=
Equivalent of Mg		
Accepted equivalent of Mg		=





## C. Tin (1.5 to 2 gm.) and HCl (concentrated).

## DATA AND COMPUTATIONS.

Number of c.c. of H produced by	gm. of Sn =
Corrected for	=
Weight of H (computed)	=
Weight of Sn for 1 of H }	=
Equivalent of Sn }	=
Accepted equivalent of Sn	=

2. *The percentage of oxygen in  $\text{KClO}_3$ ; also the volume of a given weight of O.*

APPARATUS: The same as that used in 1, except that a clean, dry, bulbed ignition tube is used instead of the s. n. test tube. The ignition tube has a cork stopper bored for glass tubing, which just passes through the cork and the rubber stopper of the flask; the ignition tube is arranged to be nearly horizontal. (Fig. 12.) Weigh the ignition tube, put into the bulb between 1 and 1.5 gm. of  $\text{KClO}_3$  and weigh again. Connect to the rest of the apparatus and with careful heating (lamp) drive off as much of the oxygen as possible; finally apply the full heat of a Bunsen burner until no more oxygen comes away. Cool, weigh tube and residue together, also measure the water as in 1.

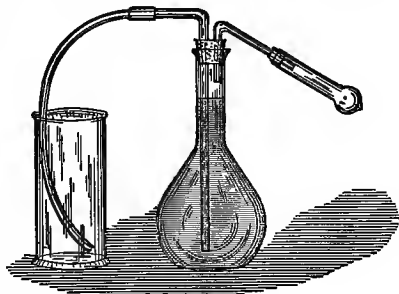


Fig. 12.

## DATA AND COMPUTATIONS.

Weight of tube +  $\text{KClO}_3$  =  
 Weight of tube = \_\_\_\_\_  
 $\text{KClO}_3$  taken =

Weight of tube +  $\text{KClO}_3$  =  
 Weight of tube +  $\text{KCl}$  = \_\_\_\_\_  
 Oxygen lost =

Percentage of O in  $\text{KClO}_3$ , by experiment =  
 Percentage of O in  $\text{KClO}_3$ , computed from formula = \_\_\_\_\_  
 Error =

Number of c.c. of O obtained from gm.  $\text{KClO}_3$  =  
 Number of c.c. of O corrected =  
 Number of c.c. of O \* computed from "oxygen lost" =  
 Number of c.c. of O computed from formula  $\text{KClO}_3$  =

\* 1 litre of O weighs 1.43 gm.

### 3. *The percentage of water of crystallization in certain salts.*

APPARATUS: Two 20c.c. porcelain crucibles may be used or two evaporating dishes of 30 to 50c.c. capacity. In this and in several similar experiments it is most convenient and accurate to weigh the dishes, and afterwards weigh the

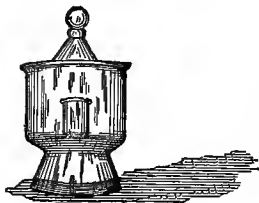


Fig. 13.

results in the dishes, and subtract the weights of the dishes. If the capacity or delicacy of the balances do not permit this, you can with sufficient care scrape the substance to be weighed from the dish with very little loss. Two of the three following should be done.

**A.** Baric chlorid. Weigh one gramme into one dish and two grammes into the other. Heat both, at first very cautiously, till all action is over. Cool, — best







in a dessicator (Fig. 13), so they will not absorb moisture, — weigh : for certainty you should heat and weigh again.

Percentage of loss from 1 gm. =

Percentage of loss from 2 gm. =

Percentage of water expressed in the formula,  $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$  =

#### B. Copper sulfate. Like A.

Percentage of loss from 1 gm. =

Percentage of loss from 2 gm. =

Percentage of water expressed by  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  =

C. Carbonate of Soda. Select perfectly fresh crystals that have not effloresced at all. Weigh rapidly. Heat with greater caution than in A and B, since the crystals contain so much water that they are likely to spatter.

Percentage of loss from 1 gm. =

Percentage of loss from 2 gm. =

Percentage of water expressed by  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  =

#### 4. *The quantity of salt in solution at a known temperature.*

Make a cold saturated solution of any of the following salts:  $\text{KCl}$ ,  $\text{NaCl}$ ,  $\text{KNO}_3$ ,  $\text{KClO}_3$ ,  $\text{Pb}(\text{NO}_3)_2$ ,  $\text{HgCl}_2$ . Use 40 to 50 c.c. of distilled water. You should make the solution the day before you wish to use it, and let it stand in a covered beaker. Now weigh two dishes (crucibles, small evaporating dishes or small beakers—see 3). Give your solution a final stirring, take its temperature, and into each dish measure with the utmost accuracy you can (with burette, pipette, or graduate) (Fig. 14) 10 c.c. of solution; weigh dishes and contents; now on the water (Fig. 15), steam or air bath, evaporate to dryness, stir and break up crystals with glass rod, rinsing rod into dish with distilled water, if necessary. A great difficulty with operations like this is

that the residue is not dry when it seems to be; so let the dishes stand in a dessicator, with concentrated  $\text{H}_2\text{SO}_4$ , over

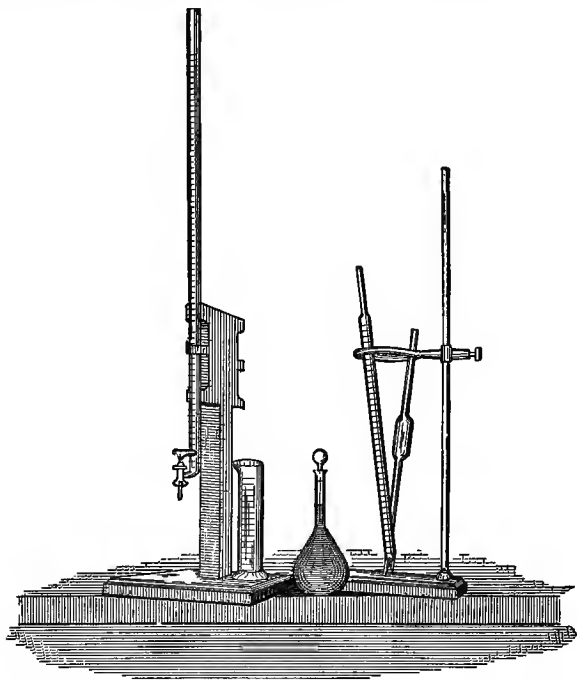


Fig. 14.

night. Weigh dishes and contents. Now you have the data to make your computations three ways.

- |   |   |
|---|---|
| <p><b>A.</b> The weight of salt<br/>temperature is</p>    | <p>in 100 parts of solution by measure at<br/>.</p>     |
| <p><b>B.</b> The weight of salt<br/>perature is</p>       | <p>in 100 parts of solution by weight at tem-<br/>.</p> |
| <p><b>C.</b> The weight of salt<br/>at temperature is</p> | <p>in 100 parts of water by measure or weight</p>       |

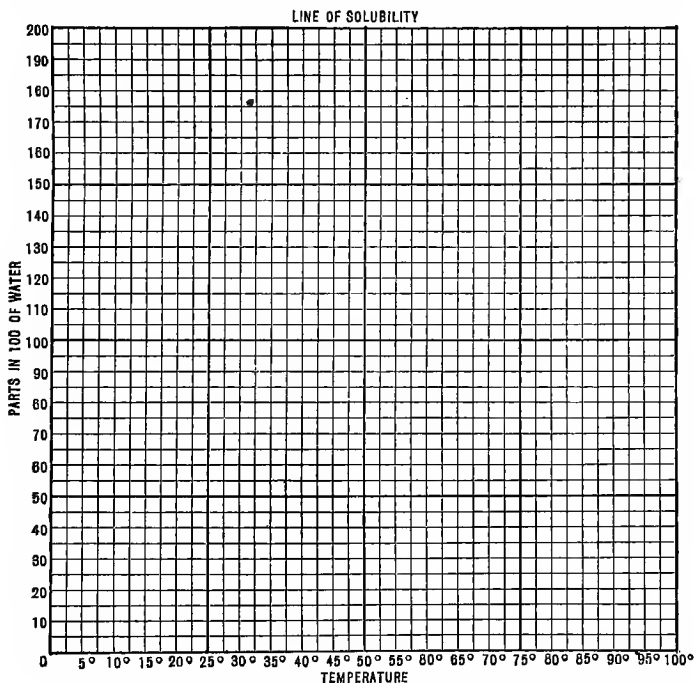




The last is the more common method of treating solutions. The first is also very important in normal solutions used in analysis.

**5.** *The solubility of a salt at different known temperatures.*

The salts and methods are similar to 4. Weigh five small dishes. Take the temperatures 20°, 30°, 40°, 50°, 60°C.



Make your saturated solution with about 60c.c. of water at 10° above the highest temperature given, and with a thermometer in the solution let it cool to 60°. Now take out,

preferably with a pipette, 10c.c. of the solution and put in one dish; weigh; let the solution cool to  $50^{\circ}$ , and take another 10c.c., etc. Of course you may select other temperatures, but you will find it difficult to manage the solution near the boiling point. Rinse your pipette with boiling water before and after using. Dry your dishes as in 4. Make your calculations as in 4, C. Then lay off the solubility in 100 parts of water at each temperature upon squared paper, and draw the curve or straight line.

In 100 of water, salt

at  $20^{\circ}$  =

at  $30^{\circ}$  =

at  $40^{\circ}$  =

at  $50^{\circ}$  =

at  $60^{\circ}$  =

### 6. *The proportion of oxygen in air, by volume.*

Take a soft glass tube  $\frac{1}{2}$  or  $\frac{3}{4}$  in. in diameter, 18 to 20 in. long; bend it like Fig. 6. (The upper end is drawn out, closed, and bent over about 1 in.; the lower end is an open loop about 2 in. long.) When the tube is cool, slip to the upper bend a piece of phosphorus half the size of a small pea. (Remember the inflammability, etc., of phosphorus.) Stand the tube in a jar of water so that the loop will fill with water; mark the height of the water in the long arm with a file or string. Return the tube to the jar, fill the jar up with water. Now gently heat the phosphorus with the lamp flame for about five minutes, taking care not to drive any air around the loop; let the tube stand in the water till it has cooled to the temperature of the room. Make the level in the long arm coincide with level in water jar; mark the height of the water. Now measure (1) contents of tube before heating; (2) contents of tube after heating. Reduce







both measurements to the same temperature, if the temperature of the room has changed sensibly during the experiment.

Air taken =  
 Residue =  
 Percentage lost (Oxygen) =

### 7. *The behavior of indicators (preliminary to 8).*

APPARATUS: Perfectly clean test tubes, distilled water, all the laboratory indicators — litmus, cochineal, lakmoid, phenol-phthalein, etc. In one test tube very dilute NaOH, in another very dilute HCl.

Now in as many test tubes as you have indicators put 5c.c. of water and add two or three drops of the various indicators. Note colors (neutral). Add one drop of HCl to each; shake thoroughly; note colors (acid); add, drop by drop, NaOH to each, shaking after each drop; note colors (alkaline); once more add HCl to note the change from alkaline to acid. (These changes are best obtained by dropping from a burette.)

NUMBER.	NAME	NEUTRAL.	ACID.	ALKALINE.	SHARP OR INDISTINCT.
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### 8. A. *The relative neutralizing power of a certain quantity, and twice that quantity of an alkali ( $\text{Na}_2\text{CO}_3$ , $10\text{H}_2\text{O}$ ) with the acids $\text{H}_2\text{SO}_4$ and $\text{HCl}$ .*

Weigh and dissolve in 50c.c. of water 2 gm. and 4 gm. of  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$ , using only perfectly fresh crystals. Add to each about 5 drops of cochineal solution (use as little as will give a perceptible color). In the most accurate measuring apparatus at hand — burette, pipette or cylinder — take

a measured quantity of quite dilute  $\text{H}_2\text{SO}_4$  (the ordinary dilute acid of the laboratory plus 3 or 4 parts of water). Now neutralize the two quantities of carbonate of soda, with constant stirring, and note the quantity of acid required for each.

2 gm.  $\text{Na}_2\text{CO}_3$  require

c.c. dilute  $\text{H}_2\text{SO}_4$ .

4 gm.  $\text{Na}_2\text{CO}_3$  require

c.c. dilute  $\text{H}_2\text{SO}_4$ .

**8. B.** Exactly the same, using quite dilute  $\text{HCl}$  instead of  $\text{H}_2\text{SO}_4$ .

2 gm.  $\text{Na}_2\text{CO}_3$  require

c.c. dilute  $\text{HCl}$ .

4 gm.  $\text{Na}_2\text{CO}_3$  require

c.c. dilute  $\text{HCl}$ .

**9.** *The relative neutralizing powers of sodium hydroxid and potassium hydroxid, with  $\text{HCl}$ .*

Weigh as rapidly as possible, in weighed beakers or watch glasses, about  $\frac{1}{2}$  in. from fresh sticks of the best  $\text{NaOH}$  and  $\text{KOH}$  obtainable. Dissolve in water and make each up to a convenient quantity—say 100 c.c. Now take of the  $\text{NaOH}$  solution enough to contain exactly .4 gm. of  $\text{NaOH}$  (quantity weighed : .4 = 100 : x. x = no. of c.c.). Color with an indicator. (In this particular case you will find phenol-phthalein very satisfactory; litmus, cochineal, lakmoid may be used.) Put in your measuring apparatus a definite quantity of dilute  $\text{HCl}$  (as in 8) and find how many c.c. are required to neutralize exactly the  $\text{NaOH}$ . Number of c.c. dilute  $\text{HCl}$  for .4 gm.  $\text{NaOH}$  = . Now measure exactly this number of c.c. of  $\text{HCl}$  into a beaker; rinse burette, first with water, then with a little of the  $\text{KOH}$  solution, and fill with the latter. Now neutralize the  $\text{HCl}$  with the  $\text{KOH}$ . No. of c.c.  $\text{KOH}$  = .

Calculate what weight of  $\text{KOH}$  was used (No. of c.c.  $\times$   $\frac{\text{KOH weighed}}{100}$ ).





The quantity of HCl which neutralizes .4 gm. NaOH, neutralizes  
KOH.

\* In what relation do these numbers stand to the formula weights of NaOH and KOH ?

**10. A.** *The quantity of sulfate, chlorid and nitrate of soda from a given weight of carbonate ; B. also by calculation the quantities of the acids required ; C. and the relations of these quantities to the formula weights.*

**A.** 1. Weigh 2.86 gm. of  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  ; dissolve in 20 to 30c.c. of distilled water in a small (weighed) evaporating dish; color with 3 or 4 drops of litmus; now add, drop by drop, dilute  $\text{H}_2\text{SO}_4$  until the liquid is pink after stirring. Evaporate to dryness upon a water (Fig. 15) or hot-air bath. Whenever during the evaporation the blue color returns to any part of the liquid, add just enough acid to restore the pink. Do not add any excess of acid. After the contents of the dish

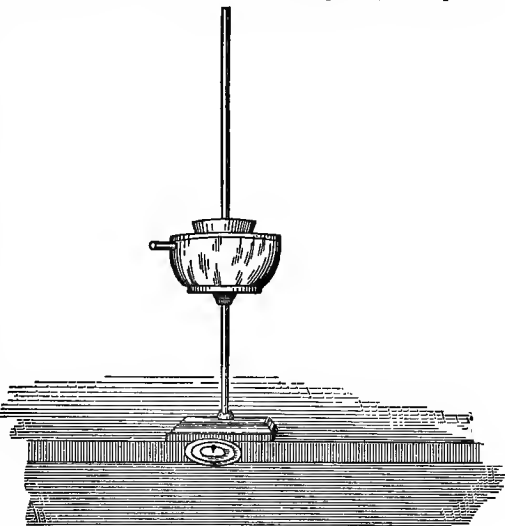


Fig. 15.

\* NOTE. — If the manipulation has been careful, the results will depend upon the purity of the hydroxids. In this laboratory the NaOH of the best makers has been found nearly pure; KOH commonly corresponds very closely to the formula  $\text{KOH}$ ,  $\text{H}_2\text{O}$ .

seem to be dry, continue the heat for an hour. Finally heat a minute or two with a Bunsen. (If the contents of the dish do not become dry on the water bath, it is because too much  $\text{H}_2\text{SO}_4$  has been added; excess can be driven out by heating in a hood, but there is danger of loss by spattering.) Cool; weigh.

Weight of  $\text{Na}_2\text{SO}_4$  obtained from 2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  =

Weight of  $\text{Na}_2\text{SO}_4$  theoretically obtainable =

B.  $\text{Na}_2\text{SO}_4$  obtained:  $\text{H}_2\text{SO}_4$  used (x) = formula weight  $\text{Na}_2\text{SO}_4$  : formula weight  $\text{H}_2\text{SO}_4$ .

Required to neutralize 2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  gm.  $\text{H}_2\text{SO}_4$ .

A. 2. Repeat this experiment with this difference: Use no indicator, and add to the solution of  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$ , 5c.c. of concentrated pure  $\text{HCl}$ , very slowly.

Weight of  $\text{NaCl}$  obtained from 2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  =

Weight of  $\text{NaCl}$  theoretically obtainable =

B.  $\text{NaCl}$  obtained:  $\text{HCl}$  used (x) = formula weight  $\text{NaCl}$  : formula weight  $\text{HCl}$ .

Required to neutralize 2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  gm.  $\text{HCl}$ .

A. 3. Repeat this as in A, 2 with 2c.c. concentrated pure  $\text{HNO}_3$ .

Weight of  $\text{NaNO}_3$  obtained from 2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  =

Weight of  $\text{NaNO}_3$  theoretically obtainable =

B.  $\text{NaNO}_3$  obtained:  $\text{HNO}_3$  used (x) = formula weight  $\text{NaNO}_3$  : formula weight  $\text{HNO}_3$ .

Tabulating: 2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  gave gm.  $\text{Na}_2\text{SO}_4$ .

gm.  $\text{NaCl}$ .

gm.  $\text{NaNO}_3$ .

2.86 gm.  $\text{Na}_2\text{CO}_3$ ,  $10\text{H}_2\text{O}$  required gm.  $\text{H}_2\text{SO}_4$ .

gm.  $\text{HCl}$ .

gm.  $\text{HNO}_3$ .

Formula weight of  $\text{H}_2\text{SO}_4$  =

$\text{HCl}$  =

$\text{HNO}_3$  =

C. How does the neutralizing power of the formula weight of  $\text{H}_2\text{SO}_4$  compare with that of the other acids, and how do they compare with each other?







**11. A.** *The weight of copper precipitated from solution by a definite weight of another metal (cadmium).*

Hold a rod of cadmium in the Bunsen flame and as it melts let the drops fall from a height of 6 or 8 in. upon a smooth surface, like stone or slate, so that they will spread out into thin disks.

Dissolve 5 gm. of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  in 50c.c. of water; add 2 drops of dilute  $\text{H}_2\text{SO}_4$ . Weigh exactly 1.12 gm. of the thin cadmium, put it in the solution and let it stand 24 hours, or until every particle of the cadmium is dissolved. Now wash the precipitated copper by decantation four times, taking care to lose no Cu, and also to pour off the water completely each time. Dry the Cu (water bath or by letting it stand in dessicator), weigh, either in the beaker, or by brushing out every particle.

Cadmium taken 1.12 gm.    Cu obtained                    gm.

Look up equivalents and atomic weights of Cd and Cu.

**11. B.** *The weight of lead precipitated by cadmium.*

The experiment is similar to **11, A.** Use 6 gm. of lead acetate, add 2 or 3 drops of acetic acid — enough to make the solution clear. Only the thinnest portion of the Cd will do for this, and before washing, you must make sure that no undissolved Cd is left.

Cd taken 1.12 gm.    Pb obtained                    gm.

**C. and D.** The same method may be applied to  $\text{AgNO}_3$  and  $\text{HgCl}_2$ . There is, however, more danger of loss of fine particles, and the weighing must be done in the beaker.

[12. *Hydrochloric acid gas is one-half hydrogen (and one-half chlorine) by measure.*

APPARATUS. See Figures. A thick glass 500c.c. Erlenmeyer filtering flask with side neck, provided with tight rubber stopper, about 1 ft.

of rubber tubing for the side neck and a stout pinchcock, which must close the tubing airtight. A jar holding 300 to 400c.c. of water. Apparatus for making HCl as below. (Fig. 16.)

Everything must be perfectly dry and the joints and stoppers must be air-tight.

Make HCl by dropping concentrated  $\text{H}_2\text{SO}_4$  from a stoppered funnel

upon rock salt in an ordinary or Erlenmeyer flask. Occasionally warm the HCl flask; now fill the side-necked flask with HCl gas by downward displacement, continuing a steady current not less than half an hour (hood). Drop 3c.c. of Mg powder (photographic) into the side-necked flask, close tight, put the free end of the rubber tube in the jar of water. Now very cautiously let water suck into the flask, to about half its capacity, by opening the pinchcock. (If the rubber tube was full of HCl, the action will be very lively, otherwise it will be slow, until the first drop of water gets over.) Close the pinchcock tight and let the flask stand until action ceases (12 hours or more).

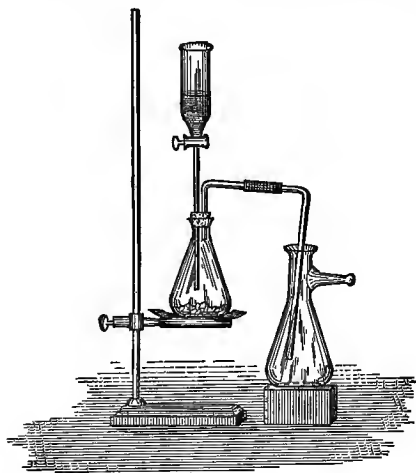


Fig. 16.





Finally tip the flask over so that the water will come into the neck (Fig. 17), open the pinchcock and let the water in the flask and jar come to equal pressure. When this condition is reached, set the flask upright again. Measure the water in it, and its total contents, making the necessary

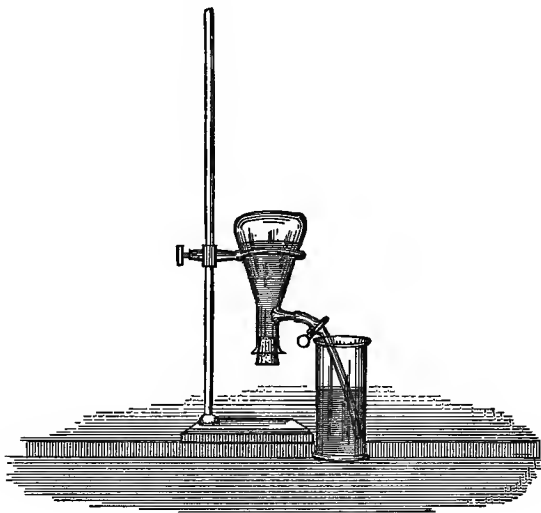


Fig. 17.

corrections. The flask was full of  $\text{HCl}$ , the  $\text{HCl}$  was absorbed by water and afterwards decomposed by  $\text{Mg}$  freeing the  $\text{H}$ . The dry gas has almost no action upon  $\text{Mg}$ .

This experiment needs very intelligent and skilful manipulation.]

Contents of flask, $\text{HCl}$	=	c.c.
Water, representing $\text{Cl}$	=	c.c.
Gas $\text{H}$	=	c.c.

**13.** *The relative volumes of CO and CO<sub>2</sub> obtained by the decomposition of ferrous oxalate.*

**APPARATUS:** Exactly that used in Experiment 2. Weigh just 1.5 gm. of well-dried FeC<sub>2</sub>O<sub>4</sub> into the ignition tube; apply gradually the full heat of a Bunsen burner; continue to heat until the residue in the tube is of uniform brownish black and all gas has ceased to come away. After cooling measure the water driven over into the jar, with the precautions and corrections previously described.

Now this experiment is to be repeated exactly, with one important difference in the conditions, viz. before beginning, dissolve 5 to 10 gm. of NaOH (or KOH) in the water of the flask; during ignition and cooling shake round the water in the flask.

Total volume of gas (CO + CO <sub>2</sub> )	=	c.c.
Volume not absorbed by NaOH (CO)	=	c.c.
Volume absorbed by NaOH (CO <sub>2</sub> )	=	c.c.

It appears that in the decomposition of FeC<sub>2</sub>O<sub>4</sub>, CO and CO<sub>2</sub> are produced in \_\_\_\_\_ volumes.

Now, assuming that the equation  $\text{FeC}_2\text{O}_4 = \text{FeO} + \text{CO} + \text{CO}_2$  is correct, which occupies the most space — 28 parts (by weight) of CO or 44 of CO<sub>2</sub>? Which occupies the most space — a certain number of molecules of CO, or the same number of molecules of CO<sub>2</sub>?

**14.** *The weight relations of certain compounds (of lead) obtained by successive transformations.*

Try to account for every slight variation from the result expected.

**A.** *The buff oxid (PbO) from white lead (2PbCO<sub>3</sub>, PbO<sub>2</sub>H<sub>2</sub>). Weigh just 10 gm. of the best, dry, white lead*







obtainable into an evaporating dish. Apply small heat and stir gently; after a time increase the heat; continue to heat and stir until the product is a powder of uniform buff color, showing no streaks of white when the little lumps are smoothed out with the rod. Cool; weigh.

10 gm. white lead give \_\_\_\_\_ gm. PbO.

If pure  $2\text{PbCO}_3$ ,  $\text{PbO}_2\text{H}_2$  theoretical yield is \_\_\_\_\_ gm. PbO.

Difference = \_\_\_\_\_

**B.** *The yellow oxid ( $\text{PbO}$ ) from the buff oxid ( $\text{PbO}$ ).* In the same dish continue to heat the product just weighed; stir, loosen from the bottom, stop heating when the contents of the dish are of uniform bright yellow; cool; weigh.

gm. buff PbO \_\_\_\_\_ give \_\_\_\_\_ gm. yellow PbO.

Theoretical yield \_\_\_\_\_ gm. yellow PbO.

Difference = \_\_\_\_\_

**C.** *The nitrate ( $\text{Pb}(\text{NO}_3)_2$ ) from the yellow oxid ( $\text{PbO}$ ).* To the weighed PbO in the dish add with stirring and gentle warming  $\text{HNO}_3$  (1 part  $\text{HNO}_3$ , 1 part water), until the yellow powder is either dissolved or changed to a white compound. Now put the dish in water or air bath and bring to dryness; break up the crystals; dry thoroughly in a dessicator; weigh.

gm. PbO \_\_\_\_\_ give \_\_\_\_\_ gm.  $\text{Pb}(\text{NO}_3)_2$ .

Theoretical yield \_\_\_\_\_ gm.  $\text{Pb}(\text{NO}_3)_2$ .

Difference = \_\_\_\_\_

**D.** *The sulfate ( $\text{PbSO}_4$ ) from the nitrate ( $\text{Pb}(\text{NO}_3)_2$ ).* Dissolve the weighed  $\text{Pb}(\text{NO}_3)_2$  in as little cold water as convenient (200 c.c. or so in a beaker); add, with constant stirring, dilute  $\text{H}_2\text{SO}_4$ , until an additional drop gives no further precipitate (let it settle before testing); now let it stand until perfectly clear above the white precipitate; pour

off the liquid, steadily but rapidly, down to the edge of the precipitate ; fill with water again and allow to settle ; decant ; repeat ; put the beaker where the precipitate will be thoroughly dried\* ; weigh.

gm. $\text{Pb}(\text{NO}_3)_2$	give	gm. $\text{PbSO}_4$ .
Theoretical yield		gm. $\text{PbSO}_4$ .
	Difference =	_____
10 gm. of white lead have given		gm. $\text{PbSO}_4$ .
Theoretical yield, direct,		gm. $\text{PbSO}_4$ .
	Difference =	_____

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\* NOTE. — A little more accurate results could be obtained by adding one-half its volume of alcohol to the water and also by decanting through a filter and brushing off the precipitate when dry. If the precipitate does not dry well, it may be moistened with alcohol and dried again.

### **SUMMARY.**

After completing as many of the experiments in Part III as seem proper to the instructor, the student is expected to prepare a complete tabulated summary and to follow this with the statement of what fact or principle he has learned from each experiment, with the reasons why absolute accuracy was not to be expected, and with any suggestions he can make for greater accuracy under the conditions of class work.





## PART IV.

### SUGGESTIONS UPON PREPARATIONS.

#### References.

1. A solution of approximate strength.
2. A solution of exact strength.
3. A saturated solution.
4. A solution of a liquid.
5. A solution of definite specific gravity.
6. The distillation of liquids.
7. The preparation and purification of gases.
8. The preparation of a solid by precipitation.
9. Crystalline salts.
10. Sublimation.

### *SUGGESTIONS UPON CHEMICAL PREPARATIONS.*

The preparation of reagents for laboratory use, chiefly in the form of water solutions, is of daily necessity and may often be made a profitable exercise for the beginner. Moreover the manufacture of many chemicals, especially of crystalline salts, is exceedingly interesting and occasionally very convenient; though it is seldom economical if the article can be purchased.

The field is very broad and only a few instructions applicable to the simpler processes can be given here. Among books that will be found useful in this field are Erdmann's "Chemischer Präparate," Curtman's "Chemical Reagents," Attfield's Chemistry, many of the larger text-

books, and also the manuals of Qualitative and Quantitative Analysis.

In this book opportunity for selection is found in the list of chemicals, Part I; the residues of many experiments collected from a number of pupils: viz. Part I, **7**, 1, **8**, 1, **11**, 4, **13**, 1, **14**, 1, **18**, 1, **19**, 1, **24**, 2, **26**, 1, **34**, 4, **35**, 2, **36**, 3, **39**, 1; Part III, **1**, **4**, **5**, **8**, **9**, **10**, **13**, **14**; larger quantities in many of the experiments given or referred to in Part I, **42-49**.

The only additional apparatus absolutely needed is evaporating dishes, holding 1 to 3 litres, funnels, about 6 in. filters, 10 to 12 in., thin glass crystallizing dishes 3 to 6 in.

**1.** *The preparation of a solution of approximate strength, of a salt in water.* Weigh the salt in grammes, measure the water in centimeters; mix; the solution may be hastened by stirring, pulverizing the salt, or by gently warming the water; whenever there is a residue of dirt or foreign material the solution is to be filtered before using. When it is destructive to filter paper, it may be decanted, syphoned off, or passed through a funnel plugged with glass wool.

When a salt is merely to be dissolved in water without specification of quantity, the novice is likely to use too little water; ten to twenty parts of water should generally be used to one of the salt.

**2.** *The preparation of a solution of exact strength.* This can only be done directly when the substance can be obtained pure and can be weighed exactly.

Weigh the substance and dissolve it in a measuring flask with a little water; after solution add water enough to make an exact measure of the whole solution. Thus  $\frac{1}{10}$  normal  $\text{AgNO}_3$  solution may be made by dissolving 17 gm. of







$\text{AgNO}_3$  in water in a litre flask and then filling the flask exactly to the mark.

(In the case of substances like liquid acids and the alkalis, not exactly weighable, solutions are made, by estimate, a little over the desired strength; their exact strength is determined by methods given in the manuals on Volumetric Analysis, then they are (or may be) diluted with water, by computation, to the exact strength desired.)

**3.** *The preparation of a saturated solution.* Merely take care that undissolved salt remains constantly at the bottom of the bottle. The solution will take care of itself, varying its strength with the variations in the temperature of the room. In the case of lime water and some very sparingly soluble substances the first water is poured off and thrown away.

**4.** *A solution of a liquid.* A measured volume of the liquid is poured into as many measured volumes of water as will give the required strength, thus: to make dilute  $\text{H}_2\text{SO}_4$  add with great caution 1 part of concentrated  $\text{H}_2\text{SO}_4$  to 4 parts of water.

**5.** *A solution of definite specific gravity.* This requires a float or hydrometer marking specific gravity (Fig. 18). (It is also done with a Westphal balance.) Thus to make  $\text{HNO}_3$  specific gravity 1.2 put the hydrometer in a jar tall enough to float it, add  $\text{HNO}_3$  till it floats, and then water till its water line is 1.2. The water should be added slowly and with stirring.

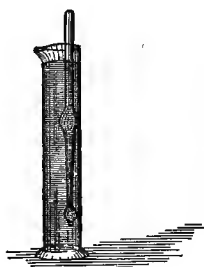


Fig. 18.

In the case of salts make a solution over the required strength and dilute.

**6.** *The preparation or purification of a liquid by distilla-*

*tion.* (See Fig. 9, also "Distillation" in larger books.) The liquid or the materials to make it are put into a side-necked round-bottomed flask, never making it more than one-half full, a thermometer passes through the cork of the flask so that its bulb is opposite the side-neck, the side-neck passes through a little cork into the inner tube of a condensing

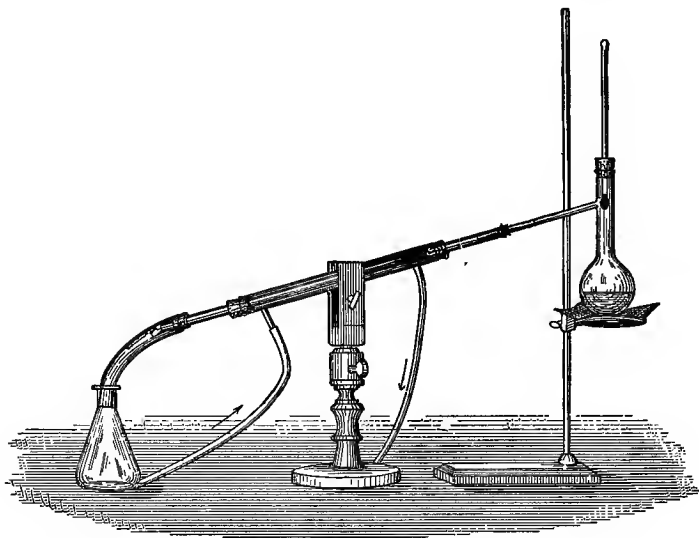


Fig. 19.

apparatus. Heat is applied through a sand bath or wire gauze, just enough to keep the liquid gently boiling; water is passed through the condenser (Fig. 19), and the distillate is collected between certain definite temperatures on the thermometer, depending on the substance. In the case of corrosive liquids, like nitric acid, a glass-stoppered flask or retort must be used. As most liquids subject to distillation





are highly inflammable this should not be undertaken without special instructions. There are many forms of distilling apparatus.

7. *The preparation and purification of gases.* Sufficient for the present purpose has been given upon preparation by the methods employed in Part I. (Let the pupils look over

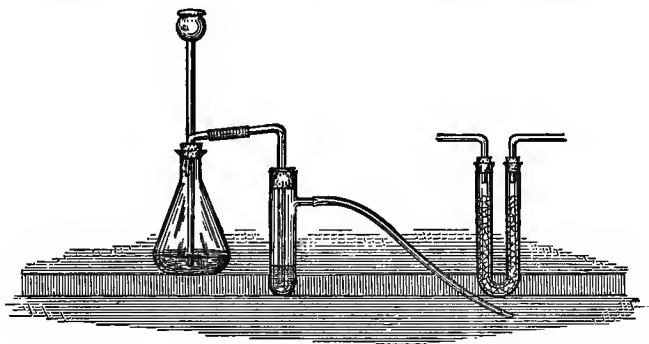


Fig. 20.

Parts I and III and describe Preparation and Collection of Gases.) The side-necked test tube is, with care, very satisfactory for small quantities of gases made either from solids or liquids; as larger quantities are required, flasks with safety tubes are much better; the Erlenmeyer flask is convenient because of its broad base. Gases soluble in water are collected by upward or downward displacement of air, over brine, or over mercury. It is often necessary to purify a gas, and this is done by causing it to pass through some substance which absorbs the impurities (Fig. 20). For instance, the gas is often washed simply by passing it through water; vapor of water is removed—the gas is dried—by concentrated  $\text{H}_2\text{SO}_4$ . A short delivery tube from the gas

generator dips under the liquid in another vessel, and a long delivery runs from this second vessel, which may be a side-necked test tube, or flask, or Wolff bottle. When the gas is to be purified by a solid the latter, in a loose and granular condition, is placed in a U-shaped tube, or in a "chlorid of calcium tube."

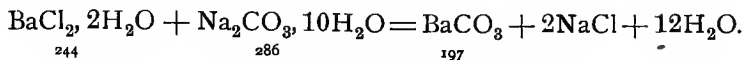
8. *The preparation of an insoluble compound by precipitation.* (See Part I, 49.) The two substances which react to produce the precipitate are in cold water solution generally. One is added to the other with constant stirring until it produces no more effect. To know that enough has been added either let the mixture settle until it is clear, or filter a few drops, and add a drop more of the reagent. It is often convenient to make an approximate computation of the quantities necessary. The liquid contains the other product of the reaction, so it is necessary to free the precipitate from this by washing. This may be done by decantation—pouring off the clear liquid down close to the precipitate and refilling with water at least five times; or by bringing the precipitate upon a filter and washing, letting each portion of water run through before adding another. Precipitates settle and filter better after thorough stirring. They may be boiled and the wash water may be hot when there is not much difference between their solubility in cold and hot water. To know the precipitate is clean the last drops of wash water are tested with some reagent for the substance being washed out; e. g. if it is a chlorid, with  $\text{AgNO}_3$ , if it is a sulfate, with  $\text{BaCl}_2$ , if it is an acid or alkali, with litmus. The next step is to dry the precipitate; this is done by keeping it in a warm place—register, steam oven, evaporating dish over water bath—and protecting it from dust. This is a slow process. Finally it is gently







separated from the filter paper. To illustrate. Prepare about 10 gm.  $\text{BaCO}_3$ .



We see that 10 gm.  $\text{BaCO}_3$  will require about 13 gm. of the chlorid, and that 15 gm. of the crystallized carbonate will be more than enough; dissolve them separately in cold water; proceed as above; common salt ( $\text{NaCl}$ ) is the product to be washed out and  $\text{BaCO}_3$  is insoluble in hot water; wash therefore with hot water until a few drops tested with  $\text{AgNO}_3$  give only the faintest cloud.

Again, make some fresh lead sulfid.  $\text{Pb}(\text{NO}_3)_2 + \text{H}_2\text{S} = \text{PbS} + 2\text{HNO}_3$ . Dissolve a gram or so of  $\text{Pb}(\text{NO}_3)_2$  in cold water; pass  $\text{H}_2\text{S}$  through the solution till the precipitate seems ready to settle in flakes and leave the liquid clear; filter and wash until the wash water is not acid to litmus. For use in **11**, 4, Part I, it is not desirable to have it dry.

**9.** *The preparation of crystalline salts.* Among the general methods of preparation are the action of acids upon metals, metallic oxids, hydroxids, and carbonates; the action of sulfuric acid upon the salts of many other acids; nearly exact double decomposition from which one of the products is insoluble and the other is soluble; direct union of chlorine with metals; the action of dry heat upon salts which give up part or the whole of their oxygen.

Each salt has individual characteristics which can only be learned by careful study.

In general crystalline salts are obtained by evaporating the water from their solutions.

If the crystals form too rapidly and imperfectly, the

evaporation has gone too far ; they should be redissolved in a little more hot water.

If they delay too long in forming, the evaporation has not gone far enough ; this can be remedied by further evaporation, or by waiting for the water to evaporate spontaneously.

Salts which are very much more soluble in hot than in cold water will often crystallize rapidly from boiling water as soon as the heat is removed.

When the size of the crystals is of little importance the solution may be evaporated until a slight crust begins to show upon the surface of the liquid ; then the heat is removed.

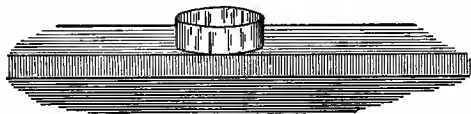


Fig. 21.

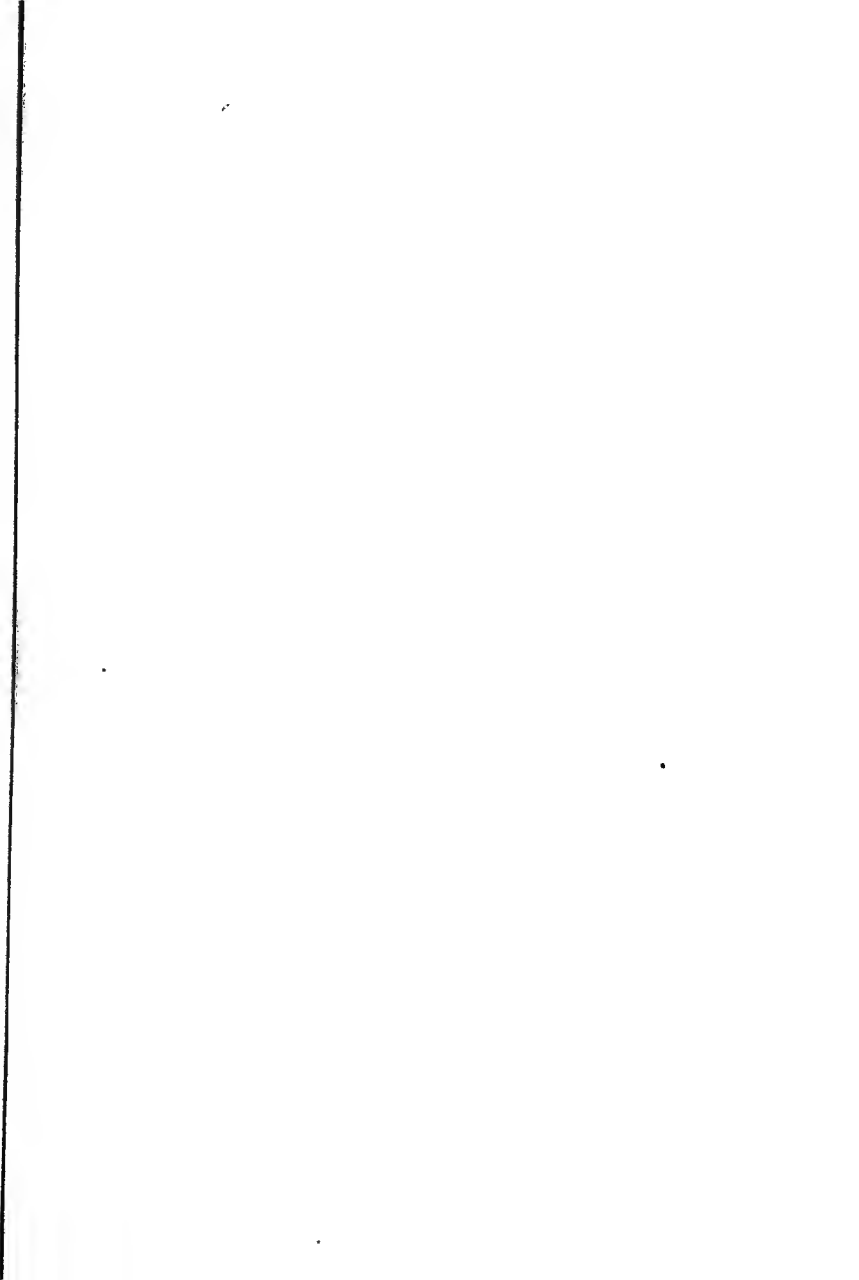
Or a drop may be put on a cool glass surface from time to time ; when it quickly shows minute crystals the heat should be removed.

Spontaneous evaporation in a warm room, or with heat insufficient to cause the liquid to boil, often results in large crystals.

Evaporation to dryness is to be avoided ; it may cause spattering, it spoils the crystals, it prevents sometimes their combination with water, it sometimes melts the salts and it leaves no mother liquor containing impurities to pour off.

Stirring or agitation of a liquid at the point of crystallization usually hastens the formation of small crystals. Sometimes this method is used to secure a crop of uniform size.

The solution at the point of crystallization is left in the evaporating dish to cool and evaporate spontaneously, or it





is poured into a crystallizing dish (Fig. 21) which has just been rinsed with boiling water.

More perfect forms may often be obtained by turning over some of the better crystals from time to time.

And in some cases very perfect crystals may be made by pouring the solution into a tall narrow jar and letting a weighted string run from the top to the bottom of the jar, not touching the sides.

On the manufacturing scale, or when the same process is to be often repeated, it is convenient to evaporate to some definite specific gravity, ascertained by a float or hydrometer.

The evaporation of small quantities of water is nicely effected under a bell jar in the presence of a beaker of concentrated  $\text{H}_2\text{SO}_4$ .

After a sufficient crop has been formed the mother liquor is poured off, the crystals are loosened with a glass rod and transferred to a funnel, or to a porcelain colander, and washed with a very little cold water and allowed to drain. The mother liquor and wash water may be evaporated for more crystals. Of course very soluble salts will not bear much washing.

After the crystals have drained well and seem fairly dry they should, if efflorescent, be bottled. Ordinary drying may be effected by spreading the crystals upon a glass plate or a plate of biscuit ware. Thorough drying even of many deliquescent salts may be effected by the bell jar and concentrated  $\text{H}_2\text{SO}_4$ .

Recrystallization consists in dissolving the crop obtained in a convenient quantity of boiling water and evaporating again.

Among the salts, products of Part I, to which these principles apply rather simply, are:  $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ ;  $\text{CuSO}_4$ ,

$5\text{H}_2\text{O}$ ;  $\text{FeSO}_4$ ,  $7\text{H}_2\text{O}$ ;  $\text{Na}_2\text{SO}_4$ ;  $\text{KCl}$ ,  $\text{NaCl}$ ,  $(\text{NH}_4)_2\text{SO}_4$ ,  $\text{NH}_4\text{NO}_3$ ,  $\text{Pb}(\text{NO}_3)_2$ .

Many double salts crystallize beautifully; e. g.  $(\text{NH}_4)_2\text{SO}_4$ ,  $\text{FeSO}_4$ ,  $6\text{H}_2\text{O}$ . Dissolve with heat 4 gm. of iron (4 three-penny nails) in  $\text{H}_2\text{SO}_4$  (40c.c. of water, 8 to 10c.c., and no more, of concentrated  $\text{H}_2\text{SO}_4$ ). After the solution is complete, add dilute  $\text{NH}_4\text{OH}$  to the liquid until a slight precipitate remains after stirring. Add dilute  $\text{H}_2\text{SO}_4$ , drop by drop, until the liquid is again acid. Filter into an evaporating dish and concentrate until a drop gives abundant crystals as it cools on a glass surface. Now transfer the hot liquid to a crystallizing dish or wide beaker, and let it crystallize. Drain off the mother liquor, break up the mass of crystals, and wash them twice with a very little cold water, then dry them.

After a very little practice the pupil will find excellent opportunity for original inquiry in seeking just the right conditions with any particular salt.

**10.** *Purification or crystallization by sublimation* is applicable to solids that can be changed to vapor by heat.

The salt is put into a dry evaporating dish without lip, or in a watch glass; a filter paper is put over the dish and an inverted funnel, or watch glass, placed above (Fig. 22). The lower dish is gently heated,—best in a sand bath,—the solid changes to vapor, passes through the paper and condenses. This is applicable to small quantities only.



Fig. 22.

Illustration:— $\text{HgI}_2$ , by precipitation and sublimation. Dissolve a drop of  $\text{Hg}$  in 2 or 3c.c. of nitro-hydrochloric acid (1 part  $\text{HNO}_3$ , 3 parts  $\text{HCl}$ ) with heat. After all has dissolved that will, dilute with 10 to 20c.c. of water, and filter into a beaker. (Solution of  $\text{HgCl}_2$  or  $\text{Hg}(\text{NO}_3)_2$  may







be used instead.) Now add with care a solution of KI. Precipitate is  $\text{HgI}_2$ , easily soluble in excess of KI. Filter, wash precipitate four or five times with very small quantities of cold water, letting each portion run through completely. Set precipitate aside to dry in the filter. Remove the dry precipitate to a watch glass, cover it with another, place on iron, and very gently heat. Notice the color changes. The  $\text{HgI}_2$  sublimes to the upper glass. Yellow crystals are rhombic; red, octahedral.

## PART V.

### CHEMICAL TERMS.

The student of chemistry must acquire early familiarity with the use of the following terms, in order to read intelligently and to speak accurately.

In many cases it must be understood that the definition is suggestive rather than complete. .

**A**—in composition; not; without; e. g. *amorphous*, not crystalline; *anhydrous*, without water.

**Acid.** A substance which neutralizes a base; sour; turns litmus red.

**Acid Salt.** A salt in which there remains displaceable hydrogen; e. g.  $\text{HNaSO}_4$ .

**Alkali.** Often applied to a base that dissolves in water.

**Alkaline.** Like an alkali; turns litmus blue.

**Allotropy.** The property of an element to exist in forms with different characteristics.

**Alloy.** Physical or chemical union of metals.

**Amalgam.** Alloy in which mercury is one of the metals.

**Amorphous.** Not crystalline.

**Analysis.** Separation into parts; applied also to the determination of the presence or quantity of substances by chemical methods.

**Anhydride.** An oxid capable of changing to an acid (or alkali) by the addition of water.  $\text{SO}_2 + \text{H}_2\text{O} = \text{H}_2\text{SO}_3$ . Sulfurous anhydride and water give sulfurous acid.

**Anhydrous.** Without water; applied to salts that have been deprived of water of crystallization; also to water-free substances in general.

—**ate.** Termination of ternary compound; implies more oxygen than —*ite*.

**Atom.** The unit of chemical change, differing for each element.

**Atomic Weight (or Mass).** The relative weight in which, or in multiples of which, an element always takes part in chemical changes.

**Base.** A substance which neutralizes an acid; usually a metal or the oxid, hydroxid, or carbonate of a metal.

**Basicity.** The number of parts (or atoms) of hydrogen in an acid which can be displaced by a metal. The basicity of  $\text{HCl}$  is 1, of  $\text{H}_2\text{SO}_4$  is 2.

**Basic Salt.** A salt having basic properties; capable of neutralizing acid.





**Bi** — Prefix meaning two, twice, double.

**Binary.** Compound consisting of two elements.

**Boil.** Changing to vapor with visible escape of bubbles.

**Bond.** The valence of an element represented by a line. H—, —O—.

**Centigrade Scale.** Freezing point at  $0^{\circ}$ , boiling point at  $100^{\circ}$ .

$$C = \frac{5}{9}(F - 32), F = \frac{9}{5}C + 32.$$

**Centimeter, cubic.** The bulk of 1 gram of water;  $\frac{1}{1000}$  litre;  $\frac{1}{4}$  teaspoonful,  $\frac{1}{30}$  fluid oz., approximately.

**Chemical Attraction.** (Chemism, Chemical Affinity.) The energy which causes and maintains chemical union.

**Combustible.** Capable of burning, usually with oxygen.

**Combustion.** Rapid chemical union with light and heat; as usually understood, rapid union with oxygen.

**Commercial.** As commonly sold; without elaborate purification.

**Compound.** The product of the union of any two or more elements; a decomposable substance.

**Concentrated.** Strong; in full strength; not weakened by dilution.

**C. P.** Chemically pure; carefully freed from foreign substances.

**Crith.** The weight of 1 litre of hydrogen — .0896 gm.

**Crystalline.** Bounded by geometrical surfaces; more or less regular and glistening.

**Decant.** To separate a liquid from a solid by careful pouring.

**Decompose.** To separate into simpler parts.

**Decrepitate.** Snapping to pieces when heated.

**Deflagration.** Burning with flashes of light, like moist gunpowder.

**Deliquescent.** Becoming liquid; applied to substances that gather moisture from the air.

**Di** — prefix, like bi —.

**Diffusion.** The intermingling of gases.

**Dilute.** Weakened by the addition of water.

**Displacement.** Collection of a gas by sending it downwards or upwards into a jar of air.

**Dissociation.** Breaking into simpler form and occupying more space.

**Distillation.** Changing a liquid to vapor by heat, condensing and collecting it again.

**Double decomposition.** The exchange of like parts between two compounds. See p. 54.

**Double Salt.** A salt containing two metals.

**Dyad.** Element with valence 2.

**Effervescence.** Escape of a gas in bubbles, usually from chemical change.

**Efflorescent.** Becoming a powder; applied to salts that lose their water of crystallization at ordinary temperature.

**Element.** One of the 72 or 73 forms of matter at present not separable into simpler parts.

**Empirical.** Experimental.

**Equation.** See pp. 14, 65, 68.

**Equivalent.** Adject., having equal chemical effect in neutralizing, precipitating, etc.; noun, the weight defined in Part II, 4, B & C.

**Evaporate.** To change to vapor; to boil away.

**Excess.** "In excess," adding a reagent in sufficient quantity to have a little left over; as in adding an acid to an alkali, add till the whole liquid is acid.

**Filtrate.** The liquid which runs through the filter.

**Formula.** The expression of a compound and its relative composition.

**Gram.** Weight of cubic centimeter of water;  $\frac{1}{1000}$  kilogram; 1000 milligrams; 15.434 grains.

**Graphical.** Representing structure by diagrams  $\boxed{\text{H}} - \boxed{\text{O}} - \boxed{\text{H}}$

**Gravimetric.** By weight.

**Halogens.** The chlorine group of elements.

**Hard.** Applied to water containing salts of lime or magnesium.

**Hydrate.** Hydroxid; better usage, a compound with  $\text{H}_2\text{O}$ .

**Hydrated.** Combined with water.

**Hydro** — Prefix applied to acids containing no oxygen;  $\text{HCl}$ , hydrochloric acid;  $\text{HClO}_3$ , chloric acid; also for hydrogen, — hydrocarbon, hydrogen and carbon compound.

**Hypo** — Prefix meaning under, less; hypochlorous acid less oxygen than chlorous acid.

**Hypothesis.** A supposition; provisional view needing proof.

—**ic.** Termination meaning higher power of combination, more of the other substance than *ous*.

—**id (ide).** Termination of name of binary compound.

**Inorganic.** Mineral; not derived from animals or plants. See *Organic*.

**Insoluble.** Not soluble — in water, unless other solvent is mentioned.

—**ite.** See *ate*.

**Kindling Temperature.** Temperature necessary to start combustion.

**Law.** The statement, based on experiment, of any of Nature's regular actions or conditions.

**Litre.** 1.06 quarts; 1000 cubic centimeters.

**Metal.** One of the heavier, lustrous, basic elements.

**Metalloid.** Non-metallic element.

**Molecule.** As commonly used, the relative weight of a compound represented







by formula; the smallest portion of a substance having the properties of the substance; a group of atoms chemically united.

**Mon** — Prefix, meaning one.

**Monad.** Element with valency one.

**Monatomic.** Supposed to have only one atom in its molecule; sometimes used for *univalent*.

**Mother liquor.** The liquid from which a salt is, or has been, crystallizing.

**Nascent.** The condition of an element at the instant of release from chemical union.

**Neutral.** Not affecting litmus; neither acid nor alkaline.

**Nomenclature.** The system of naming (chemical substances).

**Non-metal.** One of the lighter acid-forming elements.

**Normal Pressure.** 1 atmosphere; 760m.m.

**Normal Salt.** A salt in which there is no (displaceable) hydrogen.  $\text{Na}_2\text{SO}_4$ .

**Normal Solution.** See page 69.

**Normal Temperature.**  $0^\circ$  Centigrade.

**Normal Volume.** Of a gas, when measured at, or calculated to, normal pressure and temperature.

**Notation.** The system of writing (chemical substances).

**Organic.** Product, direct or indirect, from animal or vegetable life; now used of carbon compounds.

— **ous.** Termination; see — *ic*.

**Oxidize.** Cause to combine with oxygen; also to raise to higher valence.

**Oxy** —. Containing oxygen; e. g. oxysalt, oxyacid.

**Per** — Prefix, beyond; more; a greater quantity of —.

**Precipitate.** The finely divided solid formed by mixture of solutions.

**Qualitative.** Having reference only to the kind of matter.

**Quantitative.** Having reference only to the mass or weight or bulk of matter.

**Radical.** A combination of elements which frequently occurs in compounds  $\text{NH}_4$ ,  $\text{SO}_4$ .

**Reaction.** A chemical change; action with indicators which show acidity, alkalinity, neutrality; sometimes used for *equation*.

**Reagent.** Applied to the substance used to bring about a chemical change.

**Reduce.** To take away oxygen; to change to lower valence.

**Saline.** Consisting of, or belonging to, a salt.

**Salt.** The rather inactive compound resulting from a base and an acid, usually solid and crystalline.

**Saturated Compound.** In which each element employs all its chemical attraction; no place for more without either substitution or rearrangement.

**Saturated Solution.** Containing all the dissolved substance possible, at the temperature.

**Soluble.** Capable of forming solution; unless otherwise mentioned, water is the solvent.

**Solution.** A clear, uniform, liquid mixture; "homogeneous mixture which cannot be separated into constituent parts by mechanical means." *Ostwald.*

**Solution, Chemical.** In which the action of the solvent makes a new compound.

**Solution, Physical.** Applied to solution from which the original substance can be recovered by evaporation.

**Solvent.** The liquid used in making a solution.

**Specific Gravity.** Times as heavy as an equal bulk of water.

**Structural.** Attempting to show the relation of the parts in a compound.

**Sub** — Prefix, under; less; lower.

**Sublimation.** Distillation applied to a solid.

**Substitution.** One substance displacing another from a compound.

**Supersaturated.** Containing more solid than is possible under ordinary conditions.

**Supporter of Combustion.** Usually applied to oxygen.

**Suspension.** The turbid mixture of an insoluble solid with a liquid.

**Symbol.** The expression for an element and its relative weight.

**Synthesis.** Putting together to form more complex.

**Ternary.** A compound of three elements.

**Theory.** Supposition or explanation resting upon sound, but not complete, evidence.

**Tri** — Prefix, three; threefold.

**Uni** — Prefix, like mono, one.

**Valence.** The number of times 1 of H or 35.4 of Cl with which the atomic weight of an element combines. See p. 65.

**Vapor Density.** The number of times as heavy as (the same bulk of) hydrogen.

**Volumetric.** By volume or bulk measure.

**Water of Crystallization.** The definite proportion of water entering into many crystals.

—yl. The termination of the name of nearly all radicals; (OH)' hydroxyl.

Practice in, and illustration of, the correct use of the foregoing terms is a valuable exercise. For instance:—

Make correct statements illustrating the proper use of the terms *acid*, *alkali*, *amorphous*, *anhydrous*, *basicity*, *boil*, *binary*, etc.





### CHEMICAL TERMS.

Fill out the blank spaces correctly :—

Concentrated is decomposes when heated.  
is deliquescent.  
Effervescence takes place when  
is insoluble in water, etc.

## **THE REFERENCE DIAGRAM.**

The student of chemistry cannot hope to understand the theoretical side of the subject until he has acquired a sound basis of knowledge of its phenomena, laws, and terminology. He must at first accept terms in their experimental meaning. But later he must add prolonged and serious study of the science as a whole or else remain forever in a mental fog wherein premises and conclusions, phenomena and hypotheses, facts and their explanation are hopelessly mixed.

The accompanying Scheme, though not perfect, presents to the eye the whole subject and the relations of its parts.

It is intended for a twofold use :—

- (1) By advanced students as a basis for topical study ;
- (2) By students in the elementary stage for frequent reference—to avoid misconceptions and to fix the relations of one or another part to the whole. For instance, every beginner must use the notation, but a glance at the diagram shows that the notation of modern chemistry is the outgrowth of elaborate experiment and reasoning.

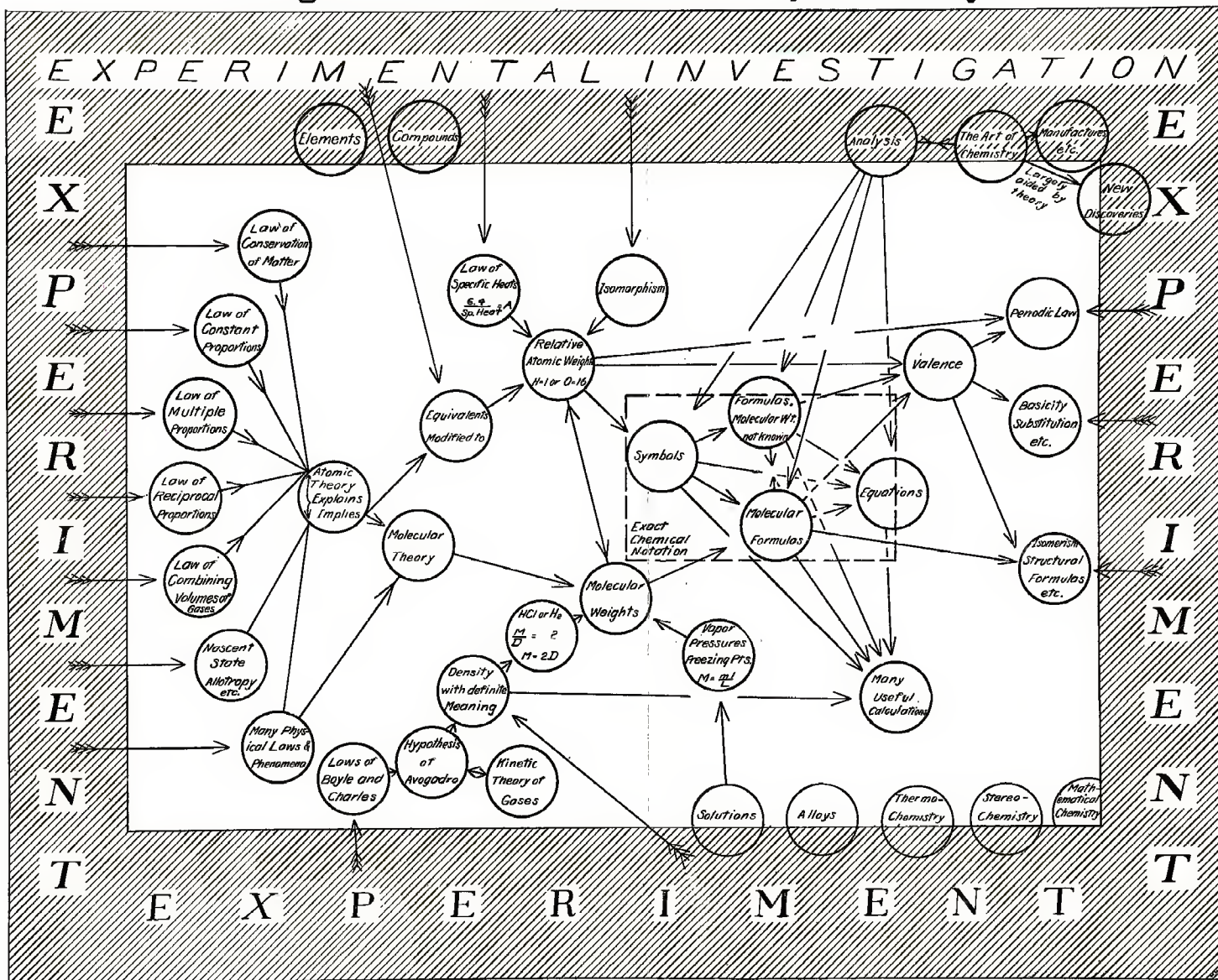
As the primary object is light upon the theories, small space is given to applied chemistry.

The diagram represents present views and is therefore only approximately historical.

In general the arrows start from *data*, *sources*, or *premises* and point to *results*, *inferences*, *conclusions*.

The arrows (with a few exceptions) point in only one direction ; but every new conception, law, discovery, fulfilled prediction which grows from these complex roots strengthens every connected part.

# Diagram for Reference and Topical Study.











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